Journal of CHEMICAL EDUCATION

OFFICIAL PUBLICATION OF THE DIVISION OF CHEMICAL EDUCATION OF THE AMERICAN CHEMICAL SOCIETY

R. D. Reed, *Chairman* N. D. Cheronis, *Vice Chairman* P. H. Fall, *Secretary* Virginia Bartow, *Treasurer*

NORRIS W. RAKESTRAW
Editor

Marilyn L. Evans *Asst. to the Editor*
Harvey F. Mack *Business Manager*
Cyrus S. Fleck *Asst. Business Manager*
B. F. Gould *Advertising Manager*
C. L. Wells *Circulation Manager*

BOARD OF PUBLICATION

Ross A. Baker, *Chairman*

S. R. Brinkley P. H. Fall
N. D. Cheronis R. D. Reed

ASSOCIATE EDITORS

Hubert N. Alyea
Princeton University
Aubison T. Burtsell
College of the City of New York
Tenney L. Davis
Mass. Institute of Technology
Ed. F. Degering
Purdue University
William A. Felsing
University of Texas
Malcolm M. Haring
University of Maryland
Philip A. Leighton
Stanford University
Laurence L. Quill
Ohio State University
Elbert C. Weaver
Bulkeley High School
Hartford, Conn.

BUSINESS & PUBLICATION OFFICES
20th and Northampton Streets
Easton, Pennsylvania

EDITORIAL OFFICE
Metcalf Chemical Laboratory
Brown University, Providence, R. I.

ADVERTISING OFFICE
11 West 42nd Street
New York, N. Y.

VOL. 18 AUGUST, 1941 NO. 8

Frontispiece—Emil Fischer IX

Editor's Outlook 351

What's Been Going On 352

Emil Fischer's Discovery of the Configuration of Glucose, A Semicentennial Retrospect, *C. S. Hudson* 353

Summer Change of Address 357

A Brief History of Thermometry, *Martin K. Barnett* 358

A Study of Some Concepts and Beliefs in Chemistry and Physics, *Lynn L. Ralya* ... 364

An Experiment in the Teaching of Inorganic and Analytical Chemistry, *William E. Morrell and Nicholas D. Cheronis* 368

Classification of Analytical Samples, *Philip J. Elving* 371

The Gram Electron, *L. F. Audrieth and M. J. Copley* 373

A Useful Chart for Qualitative Analysis, *L. P. Biefeld* 374

Demonstrations of Gas Explosions, *W. E. Thrun and J. M. Lien* 375

Solubility Generalizations, *Charles J. Heimerzheim* 377

A Simple Demonstrating Ultramicroscope, *Harold J. Abrahams and William Blitzstein* 378

An Historic Piece of Apparatus 379

A History of the Match Industry, *M. F. Crass, Jr* 380

A New Product 384

Advantages of Semimicro Technic in Teaching Qualitative Analysis, *Paul Arthur, J. A. Burrows, O. M. Smith, and Everett L. Adams* 385

The Effect of Moisture on Chemical Reactions, *George Chamlin* 386

Seventy-Five Years of Chemistry at the University of New Hampshire, *James A. Funkhouser* 389

Derivation of the Freezing-Point Equation, *Howard M. Teeter* 393

High-School Chemistry
Apparatus for Producing Liquids of Low Boiling Points, *Ernest B. Wilson* 394

Letters 395

Out of the Editor's Basket 397

Recent Books 398

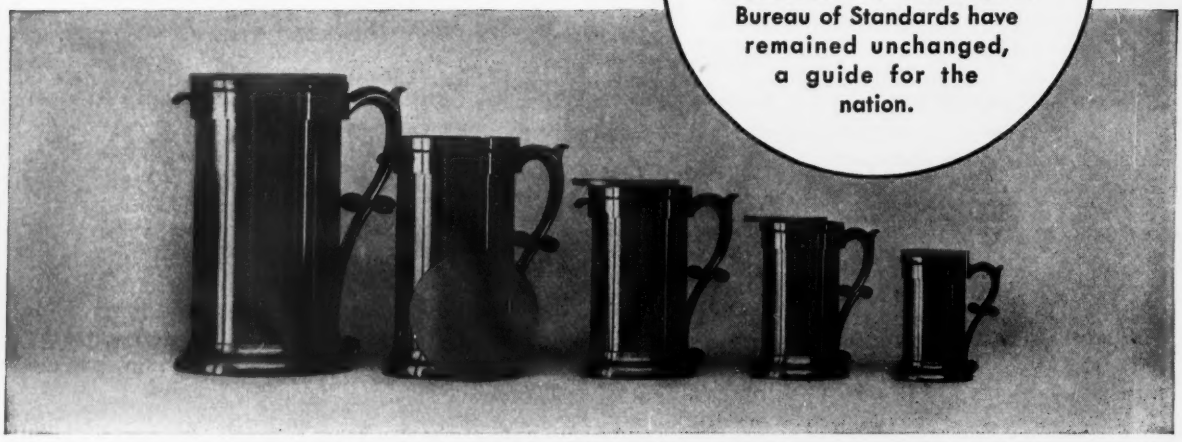
Trade Announcements XII

Subscription Price: $3.00 per year; Canada $3.50; other countries $4.00. Issued monthly. Single copies of issues in current year, 50c each; all other single copies, one dollar. Foreign remittances must be accompanied by International Money Order.
No claims for copies of journals lost in the mails can be allowed unless such claims are received within sixty (60) days of the date of issue and no claims can be allowed for issues lost as a result of insufficient notice of change of address.
Manuscript for publication should be directed to the editor. Prospective authors are urged to avoid diffuse writing and to observe the forms outlined in our Notice to Authors of Papers as it appears on page XX of the April, 1941, issue.
Entered as second-class matter, January 31, 1924, at the Post Office at Easton, Pa., under the act of March 3, 1879. Accepted for mailing at special rate of postage provided for in Section 1103, Act of October 3, 1917, authorized January 31, 1924.
The JOURNAL OF CHEMICAL EDUCATION *is indexed in the Education Index, Industrial Arts Index, and, in part, in the Quarterly Cumulative Index Medicus*

COPYRIGHT, 1941, BY DIVISION OF CHEMICAL EDUCATION, AMERICAN CHEMICAL SOCIETY

BOSTON UNIVERSITY
COLLEGE OF LIBERAL ARTS
LIBRARY

National Standards for half pint to gallon measures as furnished to the states by action of Congress in 1836. For over a century these measures of the National Bureau of Standards have remained unchanged, a guide for the nation.



FIXED STANDARDS...

Fully dependable as the National Standards for capacity measures and Mallinckrodt A.R. Chemicals. Analytical accuracy is facilitated and made more certain because of their precision. Every Mallinckrodt Analytical Reagent is made to pre-determined standards of purity with the most sensitive devices known to chemical science.

Send for new catalogue of Mallinckrodt Analytical Reagents and other chemicals for laboratory use. It contains detailed descriptions of chemicals for every type of analytical work . . . gravimetric, gasometric, colorimetric or titrimetric.

FINE CHEMICALS
Mallinckrodt
SINCE 1867

ALWAYS SPECIFY REAGENTS IN MANUFACTURER'S ORIGINAL PACKAGES

MALLINCKRODT CHEMICAL WORKS

ST. LOUIS • CHICAGO • NEW YORK • PHILADELPHIA

THE OBLIGATIONS OF A CHEMICAL LABORATORY

Every laboratory should provide the proper equipment to protect workers from accidental injury. The provision of such material helps to avoid accidents and to comply with the obligation to avoid negligence. Should accidents occur, the availability of first-aid materials suitable for treating laboratory injuries is a prime requisite.

Fisher Safety Appliances

meet the requirements for protection from accidents peculiar to laboratory operations.

for GREATER SAFETY IN THE LABORATORY—

Fisher Safety Appliances Minimize Laboratory Accidents, Avoid Negligence, Provide Protection for the Student

√ *Check List of Safety Appliances:*

Fisher First-Aid Cabinet for Laboratories, a requisite for treatment of injuries before the physician arrives. This is not an ordinary first-aid cabinet, but a special one that contains 33 first-aid items needed to treat accidents peculiar to laboratories. Supplied in metal case with an Emergency Chart outlining first-aid treatments. **Each, $14.75**

Fisher Safety Pipette Filler, completely eliminates the danger of getting harmful liquids in the mouth. A small finger-operated pump that attaches to the pipette draws up poisonous solutions, acids, alkalis, dyes, bacteriological and other harmful fluids. Liquids can be drawn into the pipette and expelled by a simple action that enables close control of the liquid level . . . **Each, $2.50**

Fire Blanket, a refuge that smothers a fire on one's clothing. It is all-wool, chemically treated, and is contained in a bright red wall case. Pulls out instantly. . . . . **Each, $20.00**

Fisher Safety Visor, protects the eyes and face while working with liquids or materials that may spatter. This transparent plastic shield can be worn with spectacles and can be tilted up out of the way. . . . . **Each, $2.50**

Fisher Safety Tongs, for handling hot beakers, flasks, casseroles and dishes. This set of four tongs should be provided by the laboratory to avoid accidents in handling hot utensils and their contents. . . . . **Per set of four, $6.00**

Fisher Safety Glass Shield, protects the student from flying glass and liquids when hazardous laboratory operations are performed. It is a stable, 30" x 14" shield that can be mounted horizontally or vertically. . . . . **Each, $16.50**

Manufactured and Distributed by

FISHER SCIENTIFIC CO.
711-723 Forbes Street • Pittsburgh, Penna.

EIMER AND AMEND
205-223 Third Avenue • New York, N. Y.

Headquarters for Laboratory Supplies

NEW 2-PURPOSE pH INDICATOR

This instrument is specialized for two purposes only—rapid, routine measurement of pH with glass electrodes, and titrations with the same electrodes.

This is our fastest, simplest, easiest-to-use and lowest-priced Indicator for glass electrodes. Its guaranteed *limit* of error of only 0.1 pH is ample for routine uses, and its strong construction makes it especially suitable for severe service.

Characteristics include:

1. Full accuracy in atmospheres of 95% relative humidity up to 85 F.
2. Easy to read because the double range of 0–8 and 6–14 pH provides a convenient overlap of 2 pH . . . a desirable feature for titrations.
3. Only 3 dial settings, including temperature compensator which ends computations, saves time, prevents errors.
4. Instrument is shielded from ordinary electrical disturbances.
5. Factory sealed and filled electrodes are highly stable.
6. Sample cup is a standard 50 ml. beaker; only 15 ml. necessary for measurement.
7. Leadwires are long enough so that electrodes can be used outside the case, for titration as well as pH.
8. The calibrated scale is longer than in any comparable pH indicator, and the pointer mechanism is outstanding for reliability and accuracy.
9. Spent batteries can't corrode rest of instrument. Maintenance is negligible.

Each indicator is sent complete with supply of chemicals and operating directions. The catalog number is 7662–A1; price, $160.00.

Speedomax Vs. The North Atlantic

Few scientists, in search of a laboratory in which to operate sensitive instruments, would pick the cabin of a coast-guard cutter. And, if they picked one, they'd probably like to tie up in a sheltered harbor.

But the Weather Bureau needs facts which can't always be gotten by working under shelter. When the nasty weather of the North Atlantic is studied, for example, someone must go where that weather is. And that "someone" often uses as a laboratory a tiny stateroom, in a Coast Guard cutter which may roll through 70 degrees, pitch like a broncho, and in general behave with unscientific violence.

This kind of work is, of course, hard on any kind of instrument, but it's especially hard on recorders. And yet the observer's task is greatly simplified if he can have at least one recorder, fast and accurate enough to jot down the temperature and other facts radioed by the stratosphere balloons which he sends up from time to time, and agile enough to operate in almost any position excepting upside down.

The Weather Bureau had had no experience with recorders for such acrobatic duty, but it set about finding such a machine. They considered the problem from every angle—accuracy, sensitivity and speed of response, as well as ability to resist *mal-de-mer*. Finally they took a Speedomax thyratron-tube recorder, set its back against a bulkhead, and welded its feet to the deck.

And that was that. Riding, as it were, the coat-tails of the Storm God, the Speedomax dutifully records the conditions which so excite that deity. In accomplishing this task, the instrument operates calmly and efficiently—requires no more attention than is needed by its fellows ashore.

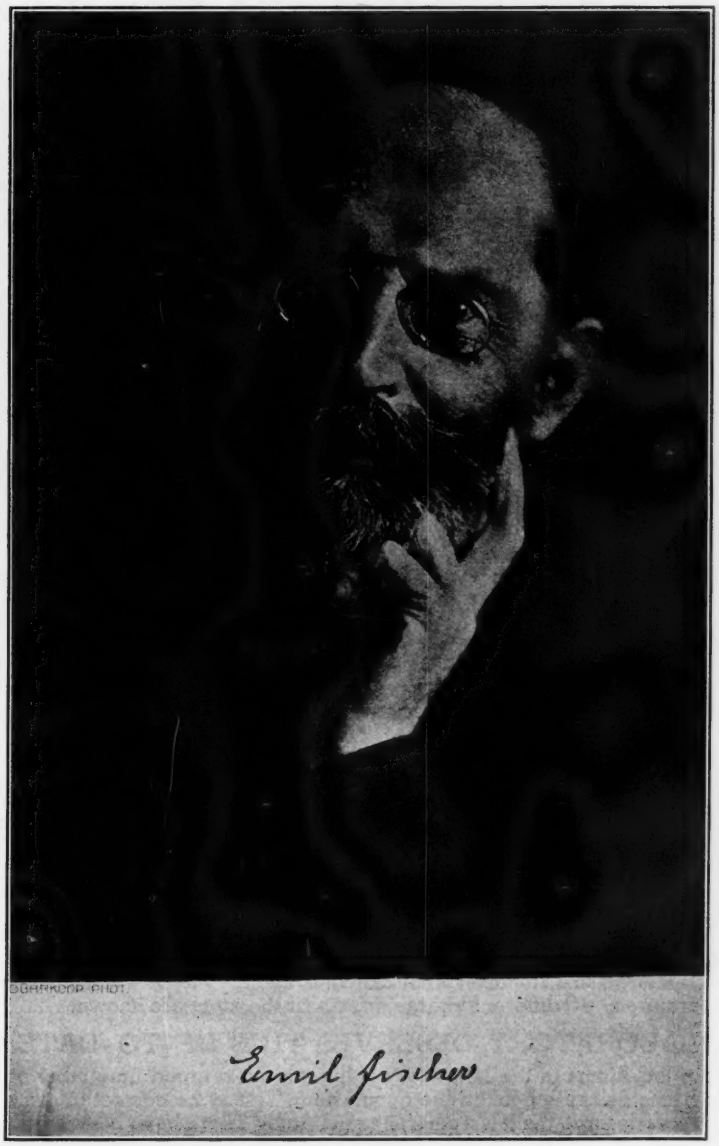
We'll be glad to send technical facts about Speedomax for a specific use. In writing, please give full details.

LEEDS & NORTHRUP COMPANY, 4976 STENTON AVE., PHILA., PA.

LEEDS & NORTHRUP

MEASURING INSTRUMENTS · TELEMETERS · AUTOMATIC CONTROLS · HEAT-TREATING FURNACES

J1 Ad E(19)

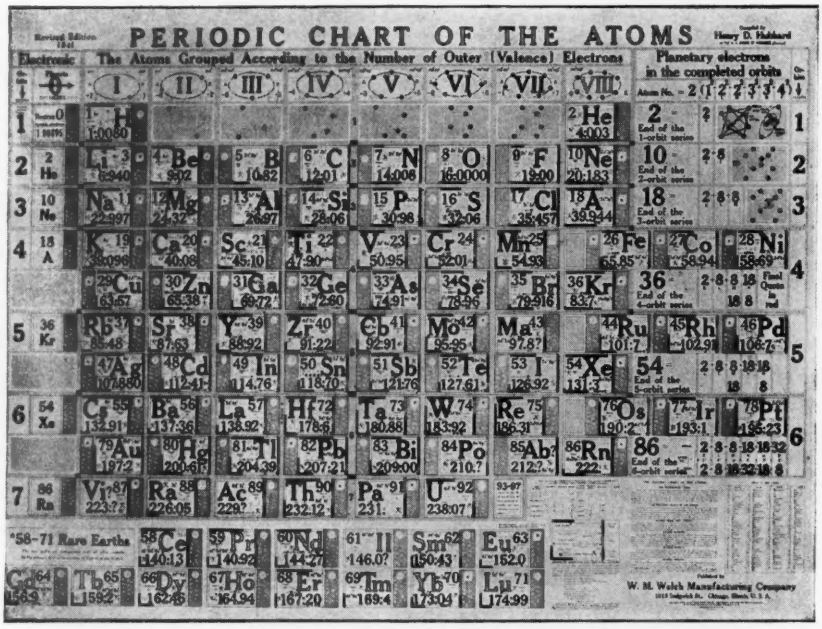


(For further information, see the article, "Emil Fischer's Discovery of the Configuration of Glucose, A Semicentennial Retrospect," by C. S. Hudson, on page 353.)

BOSTON UNIVERSITY
COLLEGE OF LIBERAL ARTS
LIBRARY

New and Up-to-the-Minute
1941 CHART OF THE ATOMS

Compiled by Henry D. Hubbard — U.S. Bureau of Standards



No. 4854

COMPLETELY REVISED—1941

This 1941 Edition represents an extensive revision under the supervision of Dr. W. F. Meggers of the U. S. Bureau of Standards. Every bit of information on the chart is brought up to this date. The latest information on the isotopes of each element appears on the chart with a new and more complete method of representing them.

LITHOGRAPHED IN 6 COLORS

The chart is lithographed in six colors and all routine information is printed in large type so as to stand out distinctly. The *atomic number* is red; the *atomic weight* is black; and by a color differentiation the large size atomic symbol indicates that the element is a solid, liquid, or gas at ordinary temperatures. *40* different characteristics of the atom are shown.

YOUR LABORATORY DESERVES THIS UP-TO-DATE CHART

An old Chart of the Atoms in the laboratory is like an old railroad time table at the station; or an out-of-date basketball rule book in the gymnasium. It may be correct in some details but not in all. Your classes deserve the latest reference material. There have been hundreds of changes since the 1924 edition.

The Chart is backed by cloth and is mounted on double split rollers top and bottom. Complete with explanatory booklet.

$5.—NO INCREASE IN PRICE—$5.

W. M. WELCH SCIENTIFIC COMPANY

Established 1880

1516 Sedgwick Street — Chicago, Illinois, U.S.A.

Editor's Outlook

IT BECOMES increasingly hard to follow the news from day to day. We are tempted to make no plans but to live solely in the present. It is a temptation to which we must not yield, of course; what counts most is what comes after. Winning a war is a life and death matter, to be sure, but after all only a means to an end. The more important, but less obvious thing to look for is an indication of what the future is to be. We must prepare for it, control it.

Not only is this the case in connection with current world affairs, but also—as always—in the field of science education. While the big cry right now is for a wide extension of technical training to fit men to carry out immediate jobs, this is no solution to the real, long-range problem: preparation for the future.

The great task of science education has always been to provide students today with an adequate basis for attack upon future problems the very nature of which is now unknown. We must master the chemistry of today not primarily for its own sake but as a preparation for that of tomorrow. In the rapidly moving world our machines of production and destruction become obsolete almost before we can install them. But if we are made of the right stuff we are not discouraged by this; it simply challenges our ingenuity and inventive capacity.

This is not only true of science's machines but of science itself. The theories of today are discarded for better ones tomorrow. Fortunately, science has within itself the seeds of survival. Change, modification, trial, error, experimentation, discard, and substitution—if you don't like these ways of getting on with things science is not for you. You can't sit in a rut for two minutes without being run over by the wheels of a new idea. You may be eligible for Social Security but don't expect to be assigned a "mental security number" which will give you license to stagnate when you feel like retiring.

The pursuit of science comes near being the fountain of eternal youth. What better way is there of staying young? You can at least keep your mind flexible even if muscles do grow stiff. And that is why we feel optimistic about scientific preparation for the future. The future will take care of itself if we can be sure of the mental flexibility which is essential to the scientific method. The success of science, in contrast to the failure of much of traditional theology, has been due to the fact that the former never shuts off argument with the statement that the last word has been said. A proper attitude on this point must be one of the outcomes of a science education if it is to equip our youth to meet the future.

In the effort to keep our science curriculum forward-looking we must of course be careful to conserve the more enduring elements while ready to eliminate, at the first evidence of their inadequacy, the inevitable temporary "fads." At the moment, we can no more tell what is fad than we can predict the future of a new hypothesis.

Still, there seem to be certain general tendencies. Of these the most notable, perhaps, is the increasing importance of mathematics and general physics. This is probably the outcome of the central position they occupy—along with chemistry—in the logical pattern of the sciences.

Moreover, there are emerging certain methods and points of view which will probably have long-standing significance and use to chemists—such as thermodynamics, statistical methods, and the quantum theory. Certainly the ability to think in terms of these will be important in the chemistry of the future, judging by present trends.

Specialization and general training are not entirely incompatible. The demands of time and the extent and complexity of material to be dealt with are making it harder and harder to give a student much breadth of training, especially in laboratory technic. The future will probably see us depending to a greater extent upon an intensive study of a relatively few problems or experiments. After all, if a student learns to study and observe natural phenomena by really thorough-going work on a very few problems he is pretty sure to develop the capacity for further scientific work, if he has a good fundamental background and the adaptability which constitutes the *sine qua non*.

What's Been Going On

FIBER D, a new rayon development by du Pont, "has a permanent crimp which gives it a wool-like appearance and feel." It can be dyed to brilliant colors, and made both mothproof and fire-retardant.

At their Point Breeze, Philadelphia, refinery the Atlantic Refining Company is to construct a new unit for the production of 100-octane gasoline for use in military aviation. Egloff states in "Petroleum Chemistry Keeps in Step with Defense" that 100-octane gasoline is now being produced in the United States at the rate of about 35,000 barrels per day. In the same article, he estimates that the potential production capacity of explosives in the United States is of the order of 85 billion pounds per year.

The plastic industry in the United States is of the order of $500,000,000 a year.

Ammonium sulfamate, according to preliminary tests, is an effective herbicide or weed-killing chemical. Successful tests have been made on Canadian thistle, dogbane, goldenrod, nettle, poison ivy, ragweed, smartweed, toad flax, wild onion, wild carrot, and yarrow, usually with timothy or red clover as an associate.

Gramicidin, a new wonder chemical, which is produced by certain soil bacteria, "is the most highly effective agent yet discovered for stopping the growth of bacteria," according to Dr. W. E. Herrell of the Mayo Clinic. Since it causes hemolysis or degeneration of the red blood cells, its use is restricted to "such cavities as the sinuses, the bladder, and infected but not bleeding wounds and any local infections such as ulcers."

Diamonds, according to reports, become permanently green when subjected to bombardment in a cyclotron.

Brushes with nylon monofilaments are now serving in the textile industry in place of the Mexican tampico fiber and Oriental hog bristle.

It is estimated that the amount of rayon produced in the United States in 1940 was sufficient to provide every woman of our democracy with seven dresses.

Oilproof Thiokol synthetic rubber, a product of the Thiokol Corporation of the Dow Chemical Company, was first produced to the extent of 4000 pounds in 1930. Present production capacity of this important rubber substitute is about 6,000,000 pounds annually.

The per capita consumption of sulfuric acid in the United States is about 125 pounds per year as compared to 25 pounds of soap and 8 pounds of table salt.

Sulfamic acid adds another link to the intriguing story of synthetic organic chemistry. Although discovered as early as 1878, sulfamic acid remained a laboratory curiosity until a process for its economically feasible manufacture was developed about three years ago by du Pont. It has already become a "tonnage chemical" for use in leather tanning, dyestuff manufacture, and in the form of its ammonium salt as a fire-retardant.

Alginates, obtained from alginic acid which accounts for about 20 per cent of the weight of dry seaweed, are being used by the British in the production of a non-inflammable synthetic fiber which has good luster and strength, at a cost somewhat less than that of rayon.

The world's foremost helium plant at Amarillo, Texas, has produced 110,000,000 cubic feet of gas during the twelve years that it has been in operation. The present capacity of the plant is 24,000,000 cubic feet annually.

Writing for the *Southern Lumberman*, J. L. Sterns states that when wood is soaked one week per inch of thickness in a saturated water solution of urea (one pound of urea per pint of water), it may be bent readily at any future time by merely heating up to 175°F. "The principle of this treatment lies in the chemical reaction of urea with lignin" to give a thermoplastic. "Urea-treated lumber is more resistant to decay than untreated stock and in drying it definitely reduces the tendency to check."

The Bureau of Mines estimates that two pounds of coal, properly utilized, can do as much work as a man can do in a day.

The nomenclature committee has approved the use of Me, Et, Pr, Bu, and Am as abbreviations for the corresponding alkyl radicals, but the extension of the practice to higher radicals is discouraged.

The transmutation of mercury into both gold and platinum on an infinitesimal scale has been effected by use of the giant Harvard cyclotron or atom smasher, as reported by Sherr and Bainbridge before the American Physical Society.

The Firestone Tire and Rubber Company of Akron, Ohio, announces the production of a weatherproof, non-checking rubber, which will resist the disintegrating effect of ozone and other weathering factors.

Upon completion of the new neoprene synthetic rubber plant at Louisville, Kentucky, which is being built, financed, and operated by the du Pont Company, the production capacity of the neoprene units will rise to about 20,000 tons annually. Neoprene has been in production for only ten years, and has the distinction of being the first non-metallic substance to be placed on the Government's mandatory priority list.

The famed Rumanian oil fields, according to estimates, produce about one-third as much oil as does the state of Illinois.

The increase in demand for phthalic and maleic anhydride is thought to be the principal factor back of the million-dollar expansion program of National Aniline and Chemicals Company.

When the steel mills are operating at maximum capacity in the United States, the production of the required amount of coke yields about 30,000,000 gallons of toluene as a by-product. When units now under construction are in operation, according to Egloff, the synthetic production of toluene from gasoline will amount to 60,000,000 gallons a year.

"We have 1800 industrial laboratories and 42,000 industrial scientists in America. But this is not enough! Every American must put to work the idea-laboratory under his hat! We must expand our idea production. We must step up our think-power," according to "A Creative Offensive for America."

The two sources of energy, states Charles F. Kettering of General Motors, are "the evaporation of water by the sun, returning as rain, and eventually becoming our water power. . . and the growth of plants." From the latter came all forms of stored energy such as coal and petroleum. The world awaits some new means of effectively picking energy out of the air.

—Ed. F. Degering

Emil Fischer's Discovery of the Configuration of Glucose[1]

A Semicentennial Retrospect

C. S. HUDSON

National Institute of Health, U. S. Public Health Service, Washington, D. C.

EMIL FISCHER'S first publication on sugars was in the year 1884, when he described their phenylosazones and recognized that he had thereby in hand for the first time crystalline derivatives which are of general applicability for the identification of sugars (*1*). He soon found that glucose and fructose yield the same osazone, that galactose yields another one and that sorbose gives a third. The analyses of the three osazones proved conclusively that the four sugars are hexoses (*1, 2*). The steps of the osazone reaction were not clear at first but in 1887 he discovered the phenylhydrazones (*2*) and was able thereby to supply the needed information and write the stoichiometric equations that are in the textbooks of today. At this early stage he discovered a new hexose, our mannose, through its phenylhydrazine derivative; the gentle oxidation of mannitol with nitric acid and the testing of the resulting products with phenylhydrazine led to the isolation of the rather insoluble mannose phenylhydrazone (*2*), from which he prepared the new hexose and named it mannose (*3*). The discovery of mannose proved of vital importance for the development of Fischer's further sugar researches: it led him to the synthesis of the substances of the mannitol group (*4*) which brought him world-wide fame; he himself has said that mannose was the key that opened the gate to this field of synthesis ["Untersuchungen," p. 11]. Mannose also led him in a direct path to his establishment of the configurations of the sugars. His study of mannose soon led to a remarkable result; glucose, fructose, and mannose yield the same phenylosazone (*5*). From this fact he drew the conclusion, now so familiar to all, that glucose and mannose are epimeric aldoses, fructose being the configurationally related 2-ketose. It is at this stage that Fischer uses for the first time the Le Bel-van't Hoff theory of the asymmetric carbon atom; he concludes that mannose is an aldehyde of mannitol, that glucose is epimeric in configuration with mannose and that consequently glucose must be the aldehyde of some alcohol other than mannitol. This last conclusion must have aroused general interest at the time because only mannitol had ever been found by the reduction of glucose with sodium amalgam. Experiment quickly supplied the decision; Meunier (*6*) and also Vincent and Delachanal (*7*) found that the true reduction product of glucose is sorbitol, mannitol being the result of a secondary reaction in the alkaline solution. Fischer verified these results and further showed that the reduction of fructose yields a mixture of mannitol and sorbitol (*8*), as would be expected from the Le Bel-van't Hoff theory.

Kiliani's discovery in 1885 of the addition of hydrocyanic acid to reducing sugars, which led him to the synthesis of acids (and their lactones) possessing one more carbon atom, was of fundamental importance in the development of Fischer's sugar researches because four years later Fischer discovered that the lactones of the sugar group can be reduced to corresponding aldoses (*10*); the synthesis of higher carbon sugars and also the reduction of the lactones of the dibasic acids of the sugar group to uronic and even to aldonic acids became possible. By this reduction he changed saccharic acid to the naturally occurring glucuronic acid (*11*). Thierfelder (*12*) had recently reduced glucuronic acid to a new and beautifully crystalline hexonic acid lactone, which he had described with precision; it is our L-gulonic lactone. Fischer showed that this lactone may be obtained in one operation by the reduction of lactonized saccharic acid, through the intermediate glucuronic acid. He and Thierfelder named the new acid gulonic acid and Fischer reduced its lactone to our L-gulose; as would be expected, the oxidation of gulonic acid yielded saccharic acid and Fischer emphasized that here for the first time had been found an example where two different hexoses (glucose and gulose) give on oxidation one and the same dibasic acid (saccharic acid), a fact which was readily interpretable (*13*) through the Le Bel-van't Hoff theory.

The story of arabinose is an interesting one. The sugar was first prepared by Scheibler (*14*) in 1868 by the acid hydrolysis of beet pulp. It was regarded as another hexose but in 1887 Kiliani (*15*) proved that it is a pentose; by hydrocyanic acid addition he converted it to a new hexonic acid and the oxidation of this yielded in turn a new six-carbon dibasic acid. It was in this same year of 1887 that Fischer discovered mannose (our D-mannose). In 1889 he prepared D-mannonic acid from mannose (*16*) by Kiliani's method of oxidation with bromine water, and in 1890 he showed (*17*) that its lactone is the optical antipode of the lactone from the hexonic acid that Kiliani had made from natural arabinose by the cyanide synthesis. Here were the D- and L-forms of mannonic lactone, and

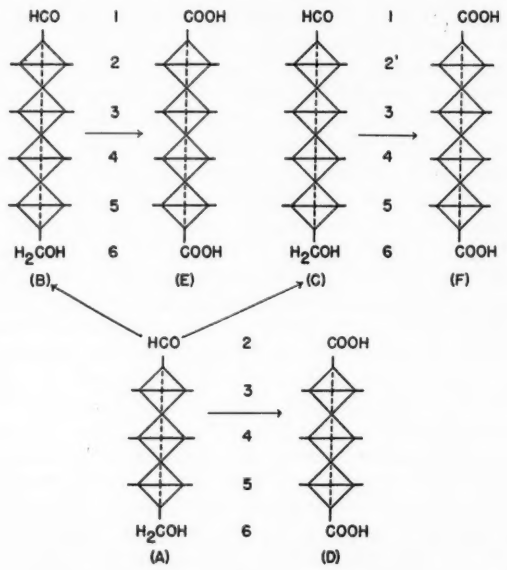
[1] Publication authorized by the Surgeon General, U. S. Public Health Service. An address delivered before the Division of Sugar Chemistry and Technology at the 101st meeting of the American Chemical Society, St. Louis, Missouri, April 10, 1941, on the occasion of the commemoration of the semicentennial anniversary of this discovery, under the honorary chairmanship of Dr. H. O. L. Fischer.

Fischer emphasized that they represent the first known pair of enantiomorphs in the sugar group. He oxidized D-mannonic acid to D-mannosaccharic acid (*18*) which proved to be the enantiomorph of the dibasic acid that Kiliani had discovered; the latter was accordingly recognized as L-mannosaccharic acid. Both acids, of course, are optically active. In 1889 Fischer and Passmore discovered the phenylhydrazides of the aldonic acids, crystalline substances which proved most useful in the separation and identification of such acids (*19*). Applying the new tool quickly, Fischer found (*20*) that natural (our L-) arabinose yields by the cyanide synthesis L-gluconic acid in addition to L-mannonic acid and that the two acids are interconvertible through heating with quinoline. These important discoveries confirmed his view that glucose and mannose are epimers. The reduction product of arabinose, crystalline arabitol, had been discovered in 1887 by Kiliani (*15*) who had reported that it showed no rotatory power; Fischer (*21*) now tested it in 1891 in a solution containing borax, following an old observation of Vignon (*22*) that borax augments the rotation of mannitol. The result was decisive in showing that arabitol is an optically active alcohol (our L-arabitol).

The sugar xylose next comes into consideration. It had been discovered by Friedrich Koch (*23*) through the acid hydrolysis of wood gum in 1886, just one year before Fischer discovered mannose. In 1889 Wheeler and Tollens (*24*) prepared its osazone for the first time and the analysis indicated that xylose was probably a pentose. Fischer (*25*) undertook then a study of xylose, after making the proper arrangement with Wheeler and Tollens. In 1891 he reported (*21*) that xylose yields by the cyanhydrin synthesis a hexonic acid which we now designate D-gulonic acid, and that this acid passes on oxidation to L-saccharic acid. He also found that the alcohol which results from the reduction of xylose, namely xylitol, shows no rotation even in the presence of borax. Arabitol was undoubtedly active, xylitol appeared to be inactive. But xylitol was a sirup; he had indeed purified it through its crystalline benzylidene derivative but the matter was of such importance that he desired further evidence. He therefore prepared the crystalline trihydroxy-glutaric acids from arabinose and xylose, respectively; the one from arabinose was active, the one from xylose was inactive. He then published his proof (*26*) of the configurations of these various sugars and derivatives, the proof that we commemorate today. The following is the introductory sentence of his proof. "*All previous observations in the sugar group are in such complete agreement with the theory of the asymmetric carbon atom that the use of this theory as a basis for the classification of these substances seems justifiable.*" To me, this sentence announces the formal advent of a theoretical chemistry of the carbohydrates.

I shall now present this proof, but not quite in its original form. Conventions of expression in this field have changed during the past fifty years and I think that the use of modified symbols and a different order for presenting the steps of the proof are in the interest of clearness. In introduction to the proof, it should be stated that it was well known at the time that the carbon chain of mannitol is a normal one because the substance yields normal hexyl iodide by reduction with hydriodic acid. It was also known that D-glucoheptonic acid yields normal heptylic acid by hydriodic reduction. The chains of glucose and mannose were thus known to be normal.

Write the formulas for a pentose and the two hexoses which it yields by the Fischer-Kiliani cyanhydrin synthesis as shown in the accompanying diagram, using Fischer's convention that the asymmetric carbon atoms (tetrahedra) have the lower edge in the plane of the paper and the corners which carry the H and OH groups lie above this plane. The arrangement of the H and OH groups is then decided through the following steps, in which the pentose is selected to be D-arabinose and in consequence the hexoses become D-glucose and D-mannose.



Step 1—By convention for the D-configurational series OH is on the right of *C*-5.

Step 2—(*D*) is optically active hence OH is on the left of *C*-3.

Step 3—D-Glucose and D-mannose are epimeric, hence the OH's on *C*-2 and *C*-2′ are opposed. Either (*B*) or (*C*) may be selected as having OH on the right, without changing the final result; here the OH is placed to the right of *C*-2 and consequently to the left of *C*-2′.

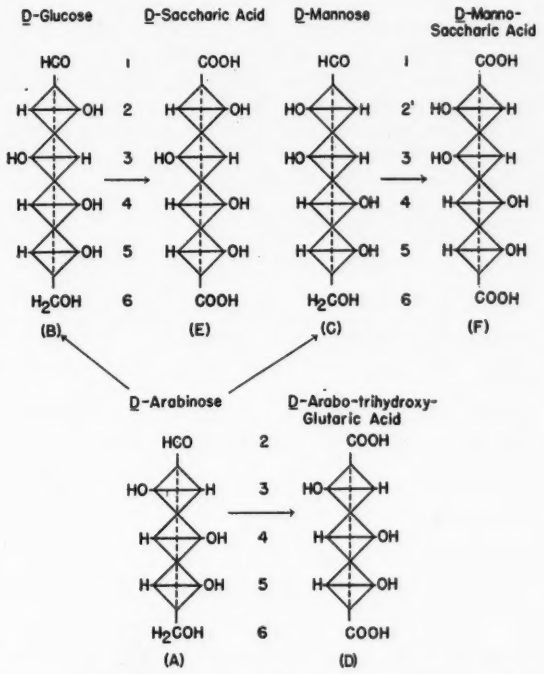
Step 4—Since both saccharic and manno-saccharic acids are optically active the configuration of neither of them can possess end-to-end symmetry; hence the OH on *C*-4 must be on the right. (If it were on the left, (*E*) would have end-to-end symmetry.)

At this stage the configuration of D-arabinose and its dibasic acid (*D*) have become established. D-Glucose and D-mannose have been limited to the configurations (*B*) and (*C*), but the correlation within this limit remains to be established. This is done by:

Step 5—Saccharic acid is obtainable from the oxidation of each of two hexoses, namely, glucose and gulose. (*E*) must therefore refer to D-saccharic acid because (*F*) cannot result from the oxidation of two hexoses. Hence (*B*)

refers to D-glucose, (*C*) to D-mannose, and (*F*) to D-manno-saccharic acid.

The proof is now complete and the formulas become:



Discoveries which have been made subsequent to 1891, several of them by Fischer himself, are the basis of various modifications of Fischer's proof. It is rare that one finds a textbook of today which follows the steps of the original proof. Possibly the preceding statement of his proof in the conventional terms of today may serve to popularize among teachers and students this classic example of exact mathematical reasoning in an experimental science.

One of the many corroborations of Fischer's proof is of some interest because of its simplicity. Referring to the third diagram, which will be understood from the preceding ones, it is noted that the inactivity of this glucoheptitol was shown by Fischer (*27*). This manno-octitol was first prepared by Fischer and Passmore (*28*); its optical inactivity was shown later (*29*).

Step 1—By convention for the D-configurational series OH is on the right of each *C*-7.

Step 2—By end-to-end symmetry *C*-2′ has OH on the right and *C*-3 has OH on the right.

Step 3—Since D-arabitol is optically active, *C*-5 must have OH on the left.

Step 4—By end-to-end symmetry *C*-4′ must have OH on the left.

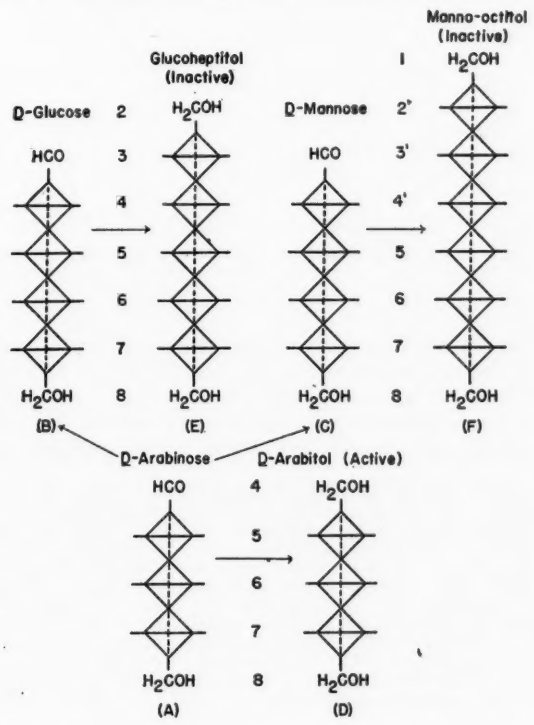
Step 5—Since glucose and mannose are epimeric, *C*-4 must have OH on the right.

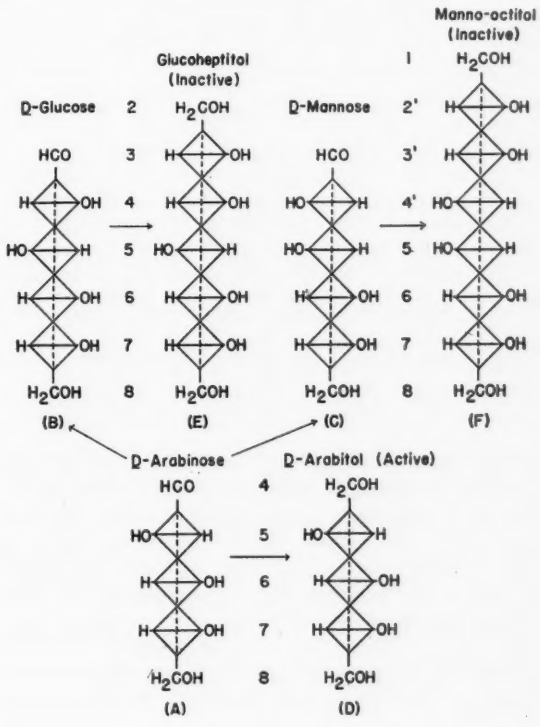
Step 6—By end-to-end symmetry *C*-6 has OH on the right.

Step 7—By end-to-end symmetry *C*-3′ has OH on the right.

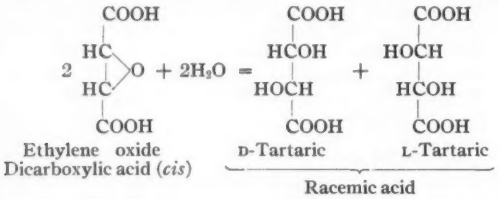
Corollary: Since carbon atoms 3′, 2′, and 1 of (*F*) have been added by cyanhydrin syntheses beginning from the carbonyl carbon 4 of (*A*), the carbons 4′, 3′, 2′, and 1 of (*F*) must represent a normal chain; hence by end-to-end symmetry the whole chain of (*F*) must be normal and therefore the chains of all the structures are normal.

When these steps have been completed the formulas become:

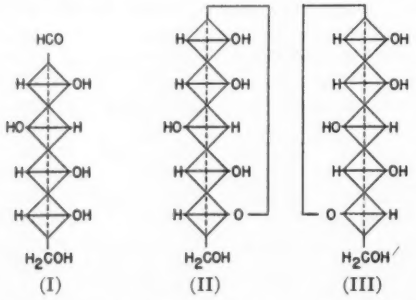




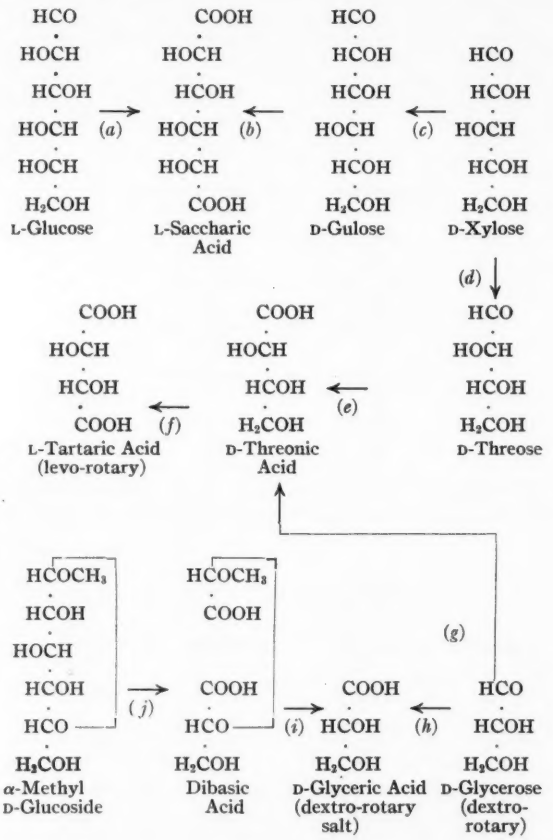
Have there arisen during these fifty years any new findings which have thrown doubt on the validity of Fischer's proof? At the end of Fischer's article of 1891 we find the statement that his arguments assume, first, that the theory of the asymmetric carbon atom is valid, and, second, that stereochemical rearrangements do not occur in the cyanhydrin synthesis of hexoses from pentoses. Five years later Paul Walden (*30*) discovered what have come to be known as the "Walden inversions." In 1908 Fischer (*31*) stated that he had been surprised that no corrections for his configurations had arisen, and attributed this to what he termed the "lucky" circumstance that his arguments had been based upon reactions which do not involve substitution on an asymmetric carbon atom. However, there arose again this question of Walden inversion through the remarkable discovery by Richard Kuhn and F. Ebel (*32*) in 1925 that the *cis* form of ethylene oxide dicarboxylic acid passes to racemic rather than to meso tartaric acid. This reaction is:



We know now that the sugars are mostly cyclic structures, and Kuhn and Ebel raised the question whether the opening of the rings at the oxygen atom may not be accompanied by a stereo change, such as they had found. On such a view the aldehydo form of D-glucose would indeed possess the accepted configuration (I) but the α-D-glucopyranose would not be (II), as generally believed, but rather (III), the formula that we commonly assign to β-L-idopyranose. W. Hückel



(*33*) published a critique of this view in which he called attention to the fact that the cyclic sugar structure is that of a cyclic hemi-acetal rather than an oxide and that Walden inversions have never been known to occur in the formation or breaking of acetal unions; he concluded therefore that the present cyclic sugar configurations are probably correct as commonly written. It seems to me that we have today additional experimental evidence that the old configurations are correct. The ring of cyclic forms of the sugars and their glycosides can be opened in two ways. The first way is by action on the oxygen atom of the ring; it is the way that Kuhn, Ebel, and Hückel discuss. The second way is by the breaking of the carbon chain of the ring with per-iodic acid; it is the reaction which E. L. Jackson and I (*34*) have carried out in the case of the methyl-D-glucopyranosides, for example, and by which we have shown that the carbon 5 of the pyranoside has the accepted configuration of the D-glyceric acid of the Fischer system. The matter is illustrated by the accompanying formulas.[2]



If Kuhn and Ebel's view were correct we should have obtained L-glyceric acid. The addition which we have made to Hückel's argument is that in our reactions the acetal which is hydrolyzed is not a cyclic acetal and therefore there can be no question of the opening of an

[2] The following references describe these reactions: (*a*) FISCHER, *Ber.*, **23,** 2611 (1890), ["Untersuchungen," (see ref. *1* for complete title) p. 372]; (*b*) FISCHER AND STAHEL, *Ber.*, **24,** 528 (1891), ["Untersuchungen," p. 394]; (*c*) FISCHER AND STAHEL, *ibid.*; (*d*), (*e*), (*f*) HOCKETT, *J. Am. Chem. Soc.*, **57,** 2260, 2265 (1935), which gives the earlier literature references; (*g*) WOHL AND MOMBER, *Ber.*, **50,** 455 (1917); (*h*) WOHL AND SCHELLENBERG, *Ber.*, **55,** 1404 (1922); (*i*) (*j*) JACKSON AND HUDSON, ref. *34*. The configuration of L-tartaric acid was established by FISCHER [*Ber.*, **29,** 1377 (1896)], ["Untersuchungen," p. 519] through the oxidation of the methyltetrose from natural rhamnose to yield D-tartaric acid (dextro-rotary).

oxygen ring in a *cis* or *trans* direction. There is thus today no valid criticism of the Fischer configurations as applied to the cyclic forms of the sugars.

The stereochemistry of the sugars that arose with Fischer's proof has developed amazingly during the subsequent half century; today we classify by his configurational formulas nearly all the simple and compound sugars and some of the polysaccharides, and his system has been extended to the amino acids and polypeptides, various types of the hydroxy acids, the inositol cyclic group of alcohols, most of the natural and synthetic glycosides, and at least one of the vitamins (ascorbic acid). In all this broad field there appears to be no established fact which disagrees with the Le Bel-van't Hoff theory of the asymmetric carbon atom; its use by Fischer to classify the sugars has proved so successful that it is today the guiding principle of the theoretical chemistry of the carbohydrates.

Throughout Emil Fischer's publications and addresses on carbohydrates one finds the most generous acknowledgment of assistance from the concurrent discoveries of others in the sugar field. He was especially appreciative of Kiliani's (*35*) work. Maquenne's (*36*) careful analyses of perseitol, which proved the substance to be a hepta-atomic alcohol rather than an isomer of mannitol as long believed, aroused his admiration (*37*). Tollens' (*38*) discovery that mannose occurs in nature (a polysaccharide in salep mucilage), followed shortly by Reiss's (*39*) preparation of it from the readily available vegetable ivory, was of the greatest practical value to Fischer, which he took occasion to emphasize (*40*). Thierfelder's (*41*) reduction of glucuronic lactone to gulonic lactone was very helpful to Fischer and the two joined in selecting the gulose name. Meunier's (*42*) discovery of the crystalline benzylidene derivative of sorbitol, through which the relationship of fructose to sorbitol was so rapidly established by Fischer, came at an opportune time in the development of Fischer's proof, as he has stated (*43*). With these delightful examples of friendly coöperation and courteous appreciation among the members of this noble group of sugar chemists of fifty years ago I shall close my story of one of the classic dramas of chemical science.

LITERATURE CITED

(1) FISCHER, *Ber.*, **17,** 579 (1884). ["Untersuchungen über Kohlenhydrate und Fermente," Verlag Julius Springer, Berlin, **1909,** Vol. I, p. 138.]
(2) FISCHER, *Ber.*, **20,** 821 (1887). ["Untersuchungen," *op. cit.*, p. 144.] Fischer did not include arabinose in his early osazone studies, presumably (C. S. H.) because Kiliani was investigating this sugar. It had always been believed to be a hexose but in 1887 Kiliani (ref. *15*) published the first analysis of arabinose osazone, which showed the sugar to be a pentose.
(3) FISCHER AND HIRSCHBERGER, *Ber.*, **21,** 1805 (1888). ["Untersuchungen," p. 289.]
(4) FISCHER, *Ber.*, **23,** 2114 (1890). ["Untersuchungen," p. 1.]
(5) FISCHER AND HIRSCHBERGER, *Ber.*, **22,** 265 (1889). ["Untersuchungen," p. 294.]
(6) MEUNIER, *Compt. rend.*, **111,** 49 (1890).
(7) VINCENT AND DELACHANAL, *ibid.*, **111,** 51 (1890).
(8) FISCHER, *Ber.*, **23,** 3684 (1890). ["Untersuchungen," p. 377.]
(9) KILIANI, *Ber.*, **18,** 3066 (1885); **19,** 767 (1886).
(10) FISCHER, *ibid.*, **22,** 2204 (1889). ["Untersuchungen," p. 315.]
(11) FISCHER AND PILOTY, *Ber.*, **24,** 521 (1891). ["Untersuchungen," p. 381.]
(12) THIERFELDER, *Z. physiol. Chem.*, **15,** 71 (1891).
(13) FISCHER AND STAHEL, *Ber.*, **24,** 528 (1891). ["Untersuchungen," p. 397.]
(14) SCHEIBLER, *Ber.*, **1,** 58, 108 (1868).
(15) KILIANI, *ibid.*, **20,** 341, 1233 (1887).
(16) FISCHER AND HIRSCHBERGER, *ibid.*, **22,** 3218 (1889). ["Untersuchungen," p. 311.]
(17) FISCHER, *Ber.*, **23,** 370 (1890). ["Untersuchungen," p. 330.]
(18) FISCHER, *Ber.*, **24,** 539 (1891). ["Untersuchungen," p. 401.]
(19) FISCHER AND PASSMORE, *Ber.*, **22,** 2728 (1889). "Untersuchungen," p. 222.]
(20) FISCHER, *Ber.*, **23,** 2611 (1890). ["Untersuchungen," p. 362.]
(21) FISCHER AND STAHEL, *Ber.*, **24,** 528 (1891). ["Untersuchungen," p. 389.]
(22) VIGNON, *Ann. chim. phys.*, [5], **2,** 440 (1874).
(23) KOCH, *Pharm. Ztg. für Russland*, **25,** 619 (1886). This journal is in the Library of the Surgeon General of the Army, Washington, D. C.
(24) WHEELER AND TOLLENS, *Ann.*, **254,** 304 (1889).
(25) FISCHER, *Ber.*, **23,** 2625 (1890). ["Untersuchungen," p. 329.]
(26) FISCHER, *Ber.*, **24,** 1836, 2683 (1891). ["Untersuchungen," pp. 417, 427.]
(27) FISCHER, *Ann.*, **270,** 64 (1892). ["Untersuchungen," p. 605.]
(28) FISCHER AND PASSMORE, *Ber.*, **23,** 2226 (1890). ["Untersuchungen," p. 579.]
(29) HANN, MACLAY, KNAUF, AND HUDSON, *J. Am. Chem. Soc.*, **61,** 1268 (1939). Compare PEIRCE, *J. Biol. Chem.*, **23,** 327 (1915).
(30) WALDEN, *Ber.*, **29,** 133 (1896).
(31) FISCHER, "Untersuchungen," p. 81.
(32) KUHN AND EBEL, *Ber.*, **58,** 919 (1925).
(33) HÜCKEL, *Z. angew. Chem.*, **39,** 849 (1926).
(34) JACKSON AND HUDSON, *J. Am. Chem. Soc.*, **59,** 994 (1937).
(35) FISCHER, ref. *4*. ["Untersuchungen," p. 3.]
(36) MAQUENNE, *Ann. chim. phys.*, [6], **19,** 1 (1890).
(37) FISCHER AND PASSMORE, *Ber.*, 2226 (1890). ["Untersuchungen," p. 574.]
(38) TOLLENS AND GANS, *Ber.*, **21,** 2150 (1888); *Ann.*, **249,** 256 (1888).
(39) REISS, *Ber.*, **22,** 609 (1889).
(40) FISCHER AND HIRSCHBERGER, *ibid.*, **22,** 3218 (1889). ["Untersuchungen," p. 308.]
(41) THIERFELDER, ref. *12*. ["Untersuchungen," p. 381.]
(42) MEUNIER, *Compt. rend.*, **110,** 579 (1890); **111,** 52 (1890).
(43) FISCHER, *Ber.*, **27,** 3189 (1894). ["Untersuchungen," p. 38.]

SUMMER CHANGE OF ADDRESS

UNTIL September 5, all correspondence for the Editorial Office of the JOURNAL should be addressed to: JOURNAL OF CHEMICAL EDUCATION, WOODS HOLE, MASSACHUSETTS.

A Brief History of Thermometry

MARTIN K. BARNETT

National Malleable and Steel Castings Company, Cleveland, Ohio

AMONG all the properties of matter, that of extension in space is one of the first which commands our attention, and the effect of heating on this property is greatest in the case of bodies in the gaseous state. Hence, it is not surprising that the first instrument designed for the detection and measurement of hot and cold was the air thermometer, whose invention is due to Galileo (1592).

GALILEO'S AIR THERMOMETER

The form of Galileo's thermometer is shown in Figure 1. The bulb B, containing the air, was brought into contact with the body whose hotness or coldness was being investigated. The resultant contraction or expansion of the air governed the height of the colored water column, the latter being regarded as a measure of the hotness or coldness in question. Thermometers of this type were early employed by physicians as a test for the fever of a patient, and the normal temperature of the human body was probably one of the first reference temperatures used in thermometry.

FIGURE 1.—GALILEO'S AIR THERMOMETER

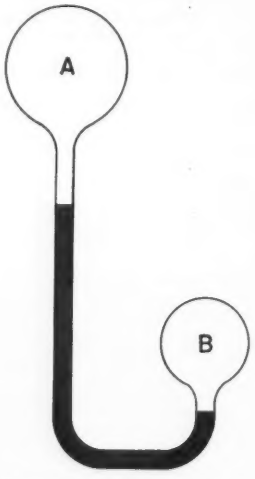


FIGURE 2.—STURM'S THERMOMETER

The dependence of the reading of the air thermometer on barometric pressure was eliminated by Sturm (1676), whose thermometer was essentially that of Figure 2. Hubin (1725), bringing the thermometer nearer to its modern form, evacuated the bulb A, the air in the bulb B being heated.

[1] This essay is a portion of a thesis to be submitted to the University of Cincinnati in partial fulfilment of the requirements for the Ph.D. degree.

TEMPERATURE AND PRESSURE

At the beginning of the eighteenth century, a new idea was introduced into thermometry by Amontons (1700), who constructed an air thermometer in which the air *pressure* was taken as a measure of the degree of heat (Figure 3). The bulb A was large, whereas the attached tube was small, about one millimeter in diameter. Hence, the volume of air remained practically constant on heating, and the height MN of the mercury column indicated the pressure.

Amontons reported that, regardless of the amount of air used, the height MN, when the air was brought in contact with boiling water, was always four-thirds of its height for cold (probably freezing) water. Hence he considered the boiling temperature of water as invariable, a fact which was not generally recognized at that time. He regarded the pressure of the air as a measure of its temperature, and thought that the "greatest cold" must correspond to zero pressure. To make his air thermometer applicable to higher temperatures, Amontons determined the pressure for the temperature of melted tallow, the remaining points of the scale being determined by interpolation and extrapolation. He pointed out the small comparability of the alcohol thermometers then in use, and undertook to graduate them by comparison with his air thermometer.

A century later the air pressure thermometer was used extensively by Lambert (1799), who introduced the variable scale for variable barometric pressures. He also regarded an "absolute cold-point" as corresponding to zero pressure.

Finally Regnault (1847), after unsuccessful attempts to make precise measurements with a constant-pressure air thermometer, adopted the constant-volume air thermometer (Centigrade scale) as his standard. It is this thermometer, with air replaced by hydrogen, which provides us with our standard thermometric scale today.

EARLY LIQUID THERMOMETERS

The relatively large coefficient of expansion of air caused the early invention of the air thermometer, but the dependence of readings on barometric pressure, along with the difficulty of manipulation, no doubt prevented it from coming into general use. Instead, liquid thermometers were used almost exclusively. Galileo had already declared expansion on heating to be a general property of matter, and Sedgwick and Tyler[2] accredit him with the introduction of the alcohol

[2] SEDGWICK AND TYLER, "A short history of science," The Macmillan Co., New York City, **1925**, p. 267.

thermometer. However, Burckhardt[3] thinks the first liquid thermometer was that of the French physician Rey (1641), who is said to have used water as thermometric substance, whereas Viviani[3] accredits it to the Grand Duke Francis II of Toscana (1631), who is known to have constructed closed alcohol thermometers. Under the latter's influence, the members of the

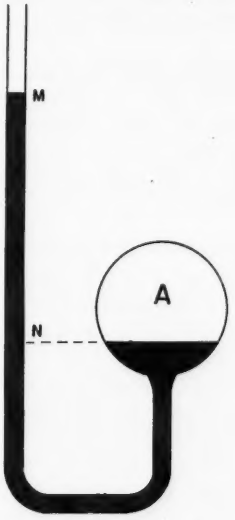


FIGURE 3.—FIRST CONSTANT-VOLUME THERMOMETER

FIGURE 4.—THE FLORENTINE THERMOMETER

Academia del Cimento at Florence came to employ such thermometers extensively, whence they were called Florentine Thermometers. The general form of this early alcohol thermometer is shown in Figure 4. It was graduated to read 20 degrees in snow and 80 degrees in "summer heat."

REFERENCE POINTS AND ARBITRARY SCALES

Boyle (1665) recommended the closed alcohol thermometer and emphasized the importance of the comparability of scales of different thermometers. He chose the freezing point of anisol as the fundamental point and thought that the numbers on the most rational scale should represent the fractional increase in volume of alcohol from the reference point, thus making more than one reference point unnecessary.

Dalence (1688) chose two reference points, the melting point of ice and the melting point of butter, and divided the intervening distance into twenty parts. Halley (1693) recommended mercury in place of alcohol as thermometric substance. Renaldini (1694) is accredited with being the first to employ the freezing and boiling points of water, together, as the two reference points. He graduated his thermometer with mixtures of freezing and boiling water in varying proportions.

[3] MACH, "Principien der Wärmelehre," 4th ed., Johann Barth, Leipzig, 1923, pp. 3–39.

Newton (1701) employed linseed oil as thermometric substance, taking for his scale numbers proportional to the expansion of linseed oil above the freezing point of water. Using this thermometer, he determined the temperatures of various fixed points.

Fahrenheit (1724), first employing alcohol, later mercury, as thermometric substance, introduced a scale which was destined to come into wide usage. He designated the temperature of a mixture of water, ice, and sal ammoniac as 0, that of melting ice as 32, and that of the blood as 96.

Reaumur (1730) reverted to Renaldini's employment of the freezing and boiling points of water as fixed points for the alcohol thermometer, dividing the intervening distance into eighty parts. Deluc, retaining the Reaumur scale, replaced the alcohol by mercury. Celsius (1742), also employing the freezing and boiling points of water as fixed points, divided this distance into 100 parts and, strangely enough, designated the melting point as 100, the boiling point as 0. Finally, Stromer reversed this choice, thus obtaining a scale close to the Centigrade scale used today.

THREE PHASES OF THERMOMETRY

Three phases may be distinguished in the history of thermometry. The one, characterized by more or less blind empiricism, produced the large variety of thermometers and thermometric scales described in the previous paragraphs. The second phase, comprising essentially critical investigations, culminated temporarily in the careful comparison by Dulong and Petit (1817) of thermometric properties of different substances and the resultant recognition of the arbitrary nature of temperature scales, the gas thermometer scales alone being generally comparable; and, finally, in the establishment of the thermodynamic scale by Thomson (1848). The third phase, beginning around the middle of the nineteenth century and extending into contemporary times, comprises the conception and technical perfection of new methods, especially in pyrometry, which, however, do not appear to contribute any new ideas to the general science of thermometry.

BOYLE'S LAW

That the first two phases are not altogether chronologically distinct is evident from the fact that we must regard the critical phase as beginning with the discovery by Boyle (1662) and Mariotte (1679) that the product of the pressure and volume of a given mass of air remains constant when the temperature is constant. The simple apparatus employed to establish this law is shown in Figure 5. The pressure was increased by pouring more mercury into the side arm, the new volume of the air being readily noted. It is probable that the form of this apparatus suggested to Amontons the idea for his air pressure thermometer, as he is known to have been acquainted with the work of Mariotte.

COEFFICIENT OF THERMAL EXPANSION

The pressure-volume relationships of gases, impor-

tant as they were for the foundation of thermodynamics, were of less immediate importance to thermometry than the much later discovered law of thermal expansion of gases. The fact that the mean thermal coefficient of expansion between 0° and 100°, as recorded on Stromer's scale of the mercury thermometer, is approximately the same for all gases was probably discovered by Charles (1787), whose work, however, was unpublished. The first published account of a complete investigation of this kind was that of Gay-Lussac (1802). The singularly simple apparatus consisted of a flask equipped with a stopcock. The flask was placed in boiling water and filled with the gas in question at room pressure. The cock was closed, the flask surrounded by melting ice, the mouth placed under water, and the cock opened. The weight of the water which flowed in and the weight of water which the flask could hold gave the volumes of the gas at the two temperatures. Using this method, Gay-Lussac found 0.00375 for the mean coefficient of expansion for different gases. In a later experiment, Gay-Lussac studied the relation between changes in volume of a quantity of air contained in a horizontal tube closed at one end with a movable thread of mercury and the temperatures of the enclosed air as recorded by Stromer's scale of the mercury thermometer, thus arriving at the important discovery that the expansion of air, on heating at constant pressure, is approximately proportional to its temperature.

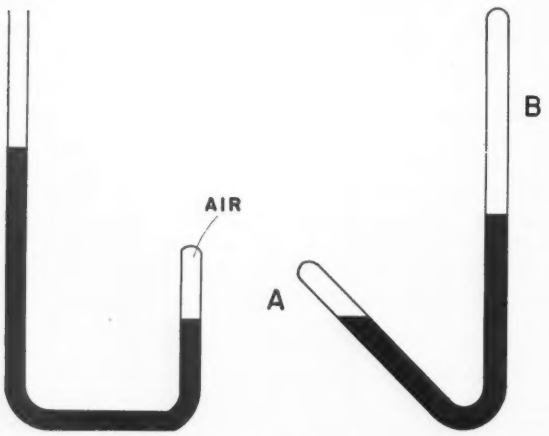


FIGURE 5.—BOYLE'S LAW APPARATUS

FIGURE 6.—DE LA TOUR'S CRITICAL POINT TUBE

DALTON'S LAW OF PARTIAL PRESSURES

In the meantime, investigations of the properties of vapors, that is, of gases in contact with the liquid phase, had begun, and these researches further emphasized the value of the temperature concept in the identification and classification of phenomena. Saussure (1783) discovered that the maximum quantity of vapor present in a space above a liquid was independent of the "nature or density of the space," but dependent on the temperature. This was out of agreement with the then prevalent view that the evaporation of a liquid was a process of solvation, the air functioning as solvent. The necessity for abandoning this point of view was substantiated by Dalton (1801) who, after finding that a given quantity of liquid, on evaporating into a vacuum, exerted the same pressure at a given temperature as when air was present with it, formulated the facts in the statement that in a space filled with gases and vapors each constituent behaves as if it alone were present.

Gay-Lussac (1802) had already discovered that ether vapor above its boiling point behaves quite like air, and this, combined with the findings of Saussure and Dalton, led to the general recognition that a vapor may exist in two states, the state of saturation and the unsaturated or superheated state.

CRITICAL PRESSURE AND TEMPERATURE

On the other hand, studies in the vapor pressure of liquids dating from those of Betancourt (1702) had established the existence of a large positive temperature coefficient, and this fact led de la Tour (1822) to suspect that at sufficiently high temperatures the density of the vapor would approach that of the liquid. He carried out experiments in a sealed tube of the type shown in Figure 6. A quantity of air at *A* was separated by an intervening body of mercury from the liquid being investigated at *B*. Upon placing the tube in a heating bath, the temperature of the liquid was given by the temperature of the bath, its pressure by the volume of the air. In this way, he arrived at values for the critical pressures and temperatures of alcohol and ether fairly close to modern values.

De la Tour's experiments probably led Faraday (1823) to the liquefaction of cyanogen and carbon dioxide. The liquefaction of other gases followed, but it was not until after Andrews (1869) had published the results of his classical investigation of carbon dioxide that the way was cleared for the liquefaction of the permanent gases by Callitet and Pictet (1877).

The key position occupied by the temperature concept in the description of phenomena had early emphasized the importance of the comparability of temperature scales. In this connection, the recognition of the maximum point in the density of water provided an important exception to Galileo's assertion that expansion with rise in temperature is a general property of matter. The maximum was recorded by Deluc (1772) as occurring at 5°C. The value now regarded as correct is 3.945°C.

RATIONALIZATION OF TEMPERATURE SCALES BY DULONG AND PETIT

For the first thorough-going investigation and comparison of thermometric scales, we are indebted to Dulong and Petit[4] (1817). Gay-Lussac had already shown that the air expansion between 0° and 100°

[4] DULONG AND PETIT, "Sur la mesure des temperatures et sur les lois de la communication de la chaleur," *Ann. chim.*, **7**, 113 (1817).

is practically proportional to the readings of the mercury thermometer (Stromer's scale). Dulong and Petit discovered marked departure from this simple proportionality at higher temperatures as the following table shows.

Hg Thermometer	*Air Thermometer*
−36	−36
0	0
100	100
360	350

Dulong and Petit did not stop here. They also determined, by an ingenious method, the heat expansions of other metals, always taking account, in their calculations, of the simultaneous expansion of the container. Considering the effect of varying the choice of thermometric substance on the expansion coefficients, they arrived at some very salient conclusions: (1) On the basis of the scale readings of the air thermometer, the expansion coefficients of other bodies increase with temperature; (2) if air be replaced by iron as thermometric substance, the expansion coefficients of other bodies decrease with increasing temperature; (3) if the pure (*i. e.*, corrected for glass expansion) volume expansion of mercury be taken as a measure of temperature, then the expansion coefficients of iron and copper increase with increasing temperature while those of platinum and air decrease.

They are thus led to emphasize that the change of volume of a body going from one state to another is an individual property of the body, and that only in the case of gases do substances conform to the same law of expansion.

Deluc and Crawford, prior to the work of Dulong and Petit, had been concerned with the question of the most scientific temperature scale and, in this connection, had sought a body whose heat expansions would be numerically equal to the quantities of heat absorbed. Dulong and Petit, while admitting that the most rational temperature scale would be one whose readings gave the heat absorbed by the thermometric substance, point out, as a result of an extensive investigation of heat capacities which they carried out for numerous substances over a wide temperature range, that such a thermometer would really have no general value or significance since the heat capacities of bodies, being peculiar to the particular substances in question, do not in general change proportionally with the same change of heat state. In other words, having defined our unit quantity of heat, it is, to be sure, perfectly possible to define a thermometer scale on which the heat capacity of a particular substance (the thermometric substance) will remain invariable with temperature; or what is the same thing, on which the scale readings will represent directly the quantities of heat absorbed by the substance. On the same scale, however, the heat capacities of other substances become variable with temperature.

The confusion prevailing in this realm of science prior to the work of Dulong and Petit is revealed in some opinions of Dalton (1808). He regarded the expansion of liquids above the freezing point as proportional to the square of the temperature increment. He thought that the vapor pressures of a liquid, corresponding to equal temperature increments, constituted a geometric progression. He speaks of the pressures of permanent gases, as well as the cooling of bodies, as following a geometric progression for equal temperature changes. On the basis of these views he proposed a temperature scale whose degree "increased in size with increasing temperature." Comparing a body to a flask containing a liquid at a definite level, Dalton stated that equal additions of heat should correspond to equal increases in temperature. At the same time he demonstrates his fundamental misunderstanding of the problem by insisting that, nevertheless, heat capacity increases with temperature! Nowhere, among all these statements, does one find Dalton attempting to define what he means by temperature.

Dalton's views on thermometry were, of course, rejected by Dulong and Petit. The latter emphasized the unavoidable dependence of the temperature scale on the arbitrary choice of thermometric substance. At the same time, they recommended the gas thermometer as superior to the others because of the general conformity of gases, in their thermal expansions, to a single simple law. Thus, the gas thermometer, having been invented first (1519) and subsequently more or less abandoned because of its practical shortcomings, was re-established, 225 years after its invention, as the most rational thermometer. The realization that it is rational in a far more fundamental sense than Dulong and Petit could understand had to await the development of thermodynamics and the subsequent discovery of the remarkable agreement between the Centigrade scale of the gas thermometer and Thomson's thermodynamic scale.

THE PERFECT GAS

As a matter of fact, the situation with respect to arbitrary scales was even less simple than Dulong and Petit believed it to be; not even the gases provided identical scales. Boyle himself recognized that his law relating pressure and volume was not exact, and this became more evident when a long series of investigations, culminating in the accurate experiments of Regnault (1847), sought to extend the law to other gases and to wider ranges of pressure. At the same time the conformity of the thermal expansion of gases to one law, advanced by Gay-Lussac, as well as the equality of the pressure and volume coefficients for a single gas, a requirement of Boyle's law, were shown to be approximated only.

Thus Regnault measured the mean volume coefficient α and mean pressure coefficient β between 0°C. and 100°C. at approximately one atmosphere pressure and obtained the following results:

	α	β
Hydrogen	0.0036613	0.0036678
Air	0.0036706	0.0036645
Carbon dioxide	0.0037099	0.0036871

To test Boyle's law directly, he measured the pv product for various pressures and discovered that air and carbon dioxide were more compressible, whereas hydrogen was less compressible, than required by the Englishman's generalization. Later Amagat (1880) found that all gases at sufficiently high pressures behave like hydrogen.

Most significant for thermometry, however, was Regnault's observation that, as the pressures under which the gases are confined are diminished, the gases conform more and more closely to the generalizations of Boyle and Gay-Lussac, the pressure and volume coefficients approaching the same value. Thus, in Regnault's work we find the basis for the important concept of the "perfect gas," constituting the ideal limit approached by real gases as they become more rarefied, and providing an ideal temperature scale approached by the real scales of actual gases. As a matter of fact, the real scales provided by hydrogen, helium, or nitrogen at ordinary pressures differ very little from the ideal scale.

REVIVAL OF AMONTONS' THERMOMETER

We have already noted how Amontons regarded the pressure of a constant volume of air as a measure of its temperature. We know from Regnault's work, just cited, that such a thermometer should at low pressures give a scale substantially identical to that of the constant-pressure type, so that preference of one over the other might seem arbitrary. Actually, however, Regnault found that the constant-pressure thermometer, because of the necessity of correcting for the temperature of the exposed gas column, was exceedingly difficult to manipulate. He was consequently led to adopt the constant-volume thermometer for accurate work. The Amontons' thermometer, as it was introduced by Regnault (Figure 7), is substantially the type employed as a fundamental standard in thermometry today.

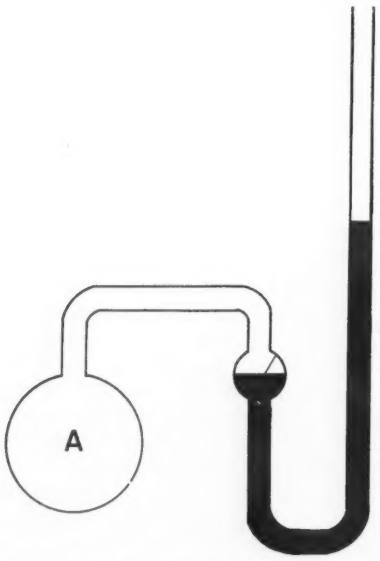


FIGURE 7.—REGNAULT'S CONSTANT-VOLUME THERMOMETER

Dulong and Petit's emphasis on the unavoidable arbitrariness in the definition of the thermometric scale also illuminated considerably the heretofore obscure problem of the scientific validity of certain methods employed in pyrometry, a term introduced by Muschenbroek (1730) for that branch of thermometry which concerns itself with the determination of high temperatures above the range where ordinary thermometers may be employed.

NEWTON'S LAW OF COOLING

One of the earliest pyrometric methods was that of Newton (1701), based on his famous cooling law. In this work Newton used a linseed oil thermometer, and as a temperature scale employed numbers proportional to the volume increase of linseed oil above the melting point of ice. He observed that when a body cools in a uniform draft of air, the rate of cooling is proportional to the excess of its temperature over that of the air, and he assumed this law to continue to hold at high temperatures incapable of direct measurement.

Thus, suppose a hot body placed in a draft of air and let $t_1, t_2, \ldots t_n, t_{n+1}$ denote times which divide the cooling period into equal intervals. If the temperature difference between body and air is $2u$ when $t = t_n$ and u when $t = t_{n+1}$, then the temperatures of the body at the times t_{n-1}, t_{n-2}, $t_{n-3}, \ldots$ will be, respectively, $4u = 2^2u$, $8u = 2^3u$, $16u = 2^4u, \ldots$. Having measured u directly, Newton was able to calculate the temperature which, assuming the law still to hold, should correspond to any previous time.

Using this method, Newton melted samples of metals and alloys by placing them on a piece of iron heated to a high temperature and then allowed the whole system to cool at a uniform rate. By observing the times at which the metals and alloys solidified, also the times at which the system attained known temperatures within the range of his linseed oil thermometer, Newton was able to calculate values for the freezing points of the metals and alloys in question, some of which were not far removed from the modern values.

CONFUSION OF TEMPERATURE AND HEAT

Considerable historical interest must be attached to Newton's attempt to prove that the scale of his linseed oil thermometer was the most rational one. He reasoned that since, in equal times, equal quantities of air particles come in contact with the iron, equally many air particles are heated by the iron by an amount proportional to the heat of the iron. Hence the "heat loss" of the iron must be proportional to the "heat" of the iron. But this loss is actually proportional to the readings of the linseed oil thermometer. Hence, says Newton, the heat of a body must be proportional to the volume increase (readings) of this thermometer.

We note from this argument that the notions of temperature and heat are by no means clearly differentiated as yet, also that the essential arbitrariness of temperature scales is not appreciated. Thus in stating that the "heat loss" of the iron is proportional to the "heat" (temperature) of the iron, Newton has already naively defined a temperature scale. Whether or not this scale agrees with some other scale, *e. g.*, that of the linseed oil thermometer, could only be determined by experiments unknown to Newton.

Dulong and Petit's later investigations revealed that considerable deviations from Newton's law of cooling occurred in ranges outside those which Newton had studied. Thus Newton's assumption of the general validity of the law was tantamount to the definition of a new temperature scale, even in the temperature range where ordinary thermometers could be employed.

BIOT'S LAW

Biot's (1804) pyrometric method was based on the following discovery. If one end of a metal bar be submitted to a constant source of heat for a sufficiently long time, and if the temperatures of the bar at distances from the heated end which form an arithmetic progression be investigated, the excesses of these temperatures over the temperature of the surroundings will be found to form a geometric progression. A century earlier, Amontons (1703) was acquainted with this phenomenon but made the mistaken assumption that the temperature t was a linear function of the distance x from the hot end, whereas Biot's law is expressed by

$$t = ar^x$$

where a is a constant and r, the ratio of the geometric progression, depends on the dimensions and nature of the metallic bar.

Biot verified the law over ordinary temperature ranges and, assuming it to hold at higher temperatures, calculated the latter from the corresponding distances. But, in modern times, we have found that r depends on t (Centigrade scale) so that Biot's law, like Newton's, is really a new definition of temperature.

BLACK'S METHOD OF MIXTURES

Another interesting pyrometric method was devised by Black (1804) on the basis of his experiments in calorimetry. If a body of mass m at a temperature u_1 is placed in a mass M of water, originally at a temperature u_2, and the water and body are allowed to assume thermal equilibrium at a temperature u, Black found that the heating of the water is proportional to $ms(u_1 - u)$, where s is a constant characteristic of the body. In fact,

$$ms(u_1 - u) = M(u - u_2)$$

or

$$u_1 = uM\,(u - u_2)/ms$$

If m and s are small and M large, u and u_2 remain within the ordinary range of temperature and can be directly determined with a thermometer. Assuming the law to hold for all heat states, *i. e.*, at all temperatures, Black calculated the very high temperature u_1 of, *e. g.*, a small piece of metal, by introducing it into the calorimeter.

But, as we have seen, Dulong and Petit discovered that s, the specific heat, is a function of temperature, as ordinarily defined, so that Black's method, like those of Newton and Biot, was tantamount to the definition of a new temperature scale. (Indeed, Renaldini (1694) had already used such a definition when he graduated his thermometers by the method of mixtures.)

A similar change of points of view occurred in connection with the platinum resistance pyrometer introduced by Siemens (1871). It was at first thought that the electrical resistance R_t was related to the conventional temperature t by the linear equation

$$R_t = R_0\,(1 + \alpha t)$$

where R_0 is the resistance for $t = 0$ and α a constant. Later it was discovered that a more complex relation

$$R_t = R_0\,(1 + \alpha t + \beta t^2)$$

was necessary, and the discrepancy was recognized by referring to the temperature, calculated from the linear expression, as the "platinum temperature."

THIRD PHASE OF THERMOMETRY

With the reference to the platinum pyrometer we are already well into the third phase of the history of thermometry, that phase characterized by technical improvements and increased precision, also by the introduction of new methods.[5] An electrical method developed prior to the resistance pyrometer was based on the measurement of thermoelectromotive force in metals. The difference in potential between the junctions of two metals maintained at different temperatures was discovered by Seebeck (1821), and the converse phenomenon by Peltier (1834). The thermodynamical theory was furnished by Thomson (1856). Becquerel (1826) made an unsuccessful attempt to employ a thermocouple for high temperature measurement. The first practical though crude device was constructed by Pouillet (1836), who used an iron-platinum couple and a low resistance tangent galvanometer. In the hands of LeChatelier (1887), the thermocouple was perfected until it became an instrument of real practical value, especially where high temperatures, small temperature differences, or the temperatures at points (*i. e.*, over very small areas) are to be determined. In general, the greater complexity of the thermocouple, also the resistance thermometer, is offset by the advantages of greater range and delicacy and by the fact that they may be employed in investigations whose circumstances make the ordinary liquid thermometers of no value, *e. g.*, determination of furnace temperatures.

Other properties have been employed as indications of temperature. These include wave length of sound,

[5] The consideration of Thomson's thermodynamic temperature scale is beyond the scope of this essay.

gaseous refractivity, rate of gaseous effusion through a small orifice, volume or pressure of solids, degree of dissociation of a substance, intensity of radiation, and magnetic moment.

All such properties, as well as the pyrometric methods previously discussed, furnish only auxiliary standards, which must be calibrated by comparison with the fundamental perfect gas (thermodynamic) scale. To be sure, the constant-volume, rather than the constant-pressure gas thermometer is, for practical reasons, taken as the reference standard. Thus, "the degree Centigrade is defined as one-hundredth part of the change in pressure of a constant volume of hydrogen gas, when its temperature is changed between that of melting pure ice and that of steam from boiling water at normal pressure, the initial pressure of the gas at 0° being 1000 millimeters of mercury."[6]

By taking into account deviations of the gases from ideal behavior, the fundamental scale may be extended so as to continue to conform to the thermodynamic scale. Whence the temperatures of certain fixed points may be determined, and these, in turn, may be employed for the calibration of the auxiliary instruments used in practice.

[6] GRIFFITHS, "Methods of measuring temperature," 2nd ed., Charles Griffin and Co., London, **1925**, p. 18.

A Study of Some Concepts and Beliefs in Chemistry and Physics

LYNN L. RALYA *The Citadel, Charleston, South Carolina*

THIS report presents the results pertaining to chemistry and physics of an investigation made to determine, for a group of entering freshmen, (1) the degree of mastery of some specific elementary and basic scientific concepts and principles; (2) the degree of prevalence of some misconceptions and unfounded beliefs hindering such mastery; and (3) the relationship between the acceptance of the correct and the rejection of the false concepts and beliefs and general scholastic aptitude (general intelligence). The investigation was motivated by the hypotheses: (1) that there are many scientific concepts, principles, and beliefs which are so relatively simple and so significant for lay thinking that they could and should become a part of the mental equipment of most people, and (2) that investigations of specific strengths and weaknesses in scientific understanding should lead, eventually, to improvement in science teaching at all educational levels comparable to the improvement in language teaching brought about by study of the relative frequency of errors in English usage or in arithmetic teaching by the determination of the relative difficulty of the different arithmetical processes.

THE SUBJECTS

The subjects coöperating in the investigation were the entering freshmen at a state college for men located in the Southeast and enrolling about 1000 students. According to comparative results on the American Psychological Examination, these students are representative in general scholastic aptitude of the freshmen entering colleges and universities in the region. A large number of the subjects had encountered either chemistry or physics in high school, although relatively few had had courses in both.

THE TESTS

Two tests were used: (1) a "Science Inventory" constructed especially for the investigation and (2) the American Council Psychological Examination.

The "Science Inventory," consisting of 232 true-false items, was constructed by a careful selection of such concepts and principles as conceivably could be universally mastered and of such basic misconceptions and unfounded beliefs as were known or suspected to exist. The sources from which selection was made included (1) state courses of study in elementary science; (2) the book, "The Story of Human Error," edited by Jastrow[1]; (3) the results of an investigation of specific achievement in high-school physics carried out by the writer[2]; (4) the results of a study of the misconceptions in science held by a group of prospective elementary teachers[3]; and (5) studies of the prevalence of unfounded beliefs and superstitions made by a number of investigators.

ADMINISTRATION OF TESTS

The Science Inventory and the American Council Psychological Examination were both administered during freshman week under standard conditions as part of the regular testing program to all freshmen entering college. Nearly all subjects completed the Science Inventory in the hour allowed. The data of

[1] JASTROW, Editor, "The story of human error," D. Appleton-Century Co., New York City, **1936**.

[2] Digests in *Sci. Education*, **22**, No. 6 (Nov., 1938); CURTIS, "Digest of investigations in the teaching of science," Vol. III, Blakiston, Philadelphia, Pa., **1939**.

[3] RALYA AND RALYA, "Some misconceptions in science held by prospective elementary teachers," *Sci. Education*, **22**, No. 5. (Oct., 1938).

this report were obtained from the papers of the 325 subjects who completed both tests under the conditions described.

TREATMENT OF DATA

For each item of the Science Inventory the per cent of incorrect responses was computed (1) for all 325 subjects, (2) for the 100 subjects scoring highest in scholastic aptitude according to the results obtained on the American Council Psychological Examination, and (3) for the 100 subjects scoring lowest in scholastic aptitude. The 84 items of the Science Inventory pertaining to chemistry and physics and the results obtained on them are presented in the table which follows.

RESULTS ON SCIENCE INVENTORY ITEMS CONCERNING CHEMISTRY AND PHYSICS

No.	Item	Per cent of subjects in error: Scholastic aptitude groups — All 325	Upper 100	Lower 100
1.	The four basic elements are fire, air, water, and earth	45	36	41
2.	There are smaller units than atoms entering into the makeup of matter	48	39	48
3.	Matter is electrical in nature	43	35	49
4.	The stars contain few, if any, elements not found on earth	43	38	42
5.	Atoms are hard, solid, perfectly spherical bodies	18	20	19
6.	When a substance burns it unites with oxygen	9	9	12
7.	An explosion of coal dust is the result of rapid burning	35	32	37
8.	Water is formed whenever most fuels are burned	40	34	44
9.	Water is an oxide	39	29	48
10.	Iron rust is heavier than the iron from which it is formed	54	51	53
11.	Air is a mixture rather than a compound	16	11	21
12.	When water evaporates it becomes air	39	21	53
13.	Any seed which will grow contains material which is alive	12	12	13
14.	Yeast cells are sometimes created by the fermentation of sweet substances	80	77	76
15.	Eels are sometimes generated from mud	9	7	14
16.	Insects are sometimes created by the decay of matter	27	19	43
17.	The starches and sugars are good sources of energy and heat for the body	20	10	30
18.	The fats of foods are good sources of energy and heat for the body	8	7	7
19.	Protein foods are needed to build and repair the body	1	0	2
20.	The white portion of milk is protein in character	39	29	51
21.	The white portion of milk and the lean portion of meat serve about the same purpose in the diet	36	31	36
22.	Fish are an especially good food for the brain	39	31	49
23.	Raw meat makes dogs fierce	27	16	40
24.	Raw meat makes men stronger than does meat thoroughly cooked	33	18	49
25.	Digestion of food begins in the mouth	21	9	25
26.	The goat normally digests metal objects	17	12	26
27.	The ostrich is capable of digesting any non-poisonous substance	39	35	49
28.	Quinine is falsely believed to be a cure for malaria	18	13	20
29.	Strong coffee is a good remedy for malaria	14	9	20
30.	Night air is the cause of several diseases	21	12	32
31.	Teeth that are carefully kept clean never decay	30	21	38
32.	The carbon dioxide in exhaled air is very poisonous	45	35	49
33.	If a mushroom blackens a silver spoon, it is poisonous	42	33	44
34.	The personality of an individual is determined almost entirely by the functioning of his glands	35	37	34
35.	Fish breathe air which is dissolved in water	19	25	22
36.	The green coloring matter (chlorophyll) in plants is essential to their growth	15	12	15
37.	Coal is made from plant material	6	0	11
38.	Some compounds made by plants and animals can be manufactured by man	13	10	15
39.	All material entering into the formation of a tree comes from the soil and water	43	36	54
40.	A body once in motion continues to move until something stops it	13	16	11
41.	It is more natural for a body to stop itself than to set itself in motion	66	57	66
42.	A uniform force acting constantly on a body tends to make the body move faster and faster	37	48	26
43.	Falling bodies tend to fall faster and faster	27	34	27
44.	The earth attracts a 10-pound body with ten times the force with which it attracts a one-pound body	57	59	59
45.	If it were not for the atmosphere, a 10-pound body would fall twice as fast as a 5-pound body	46	34	56
46.	The moon is largely responsible for tides	7	2	8
47.	A given pendulum may swing through a larger arc in the same time it requires to swing through a smaller arc	43	38	40
48.	A floating body displaces its own weight of the liquid in which it floats	11	10	16
49.	A totally submerged body is buoyed up by a force equal to the weight of the liquid which it displaces	15	12	18
50.	The water pressure at a given depth in a large lake is greater than at the same depth in a small lake	26	11	35
51.	The pressure at a given depth in a liquid is the same in all directions	22	7	31
52.	The atmosphere exerts pressure only when a wind is blowing	5	2	8
53.	The pressure of the atmosphere against a square inch at sea level averages about 15 pounds	20	15	28
54.	The weight of the atmosphere above a square inch at sea level averages about 15 pounds	44	40	46
55.	Nature never allows a vacuum to exist if there is any matter nearby to fill it	77	88	67
56.	A solid body weighs more when hot than when cold	29	14	31
57.	Ice possesses no heat	18	15	26
58.	Two bodies of the same weight and at the same temperature contain the same amount of heat	36	36	38
59.	The seasons are caused by the varying distance of the sun from the earth	79	64	89
60.	The temperature increases as we dig deeper and deeper into the earth	40	30	52
61.	If air is warmed, its capacity to hold moisture is increased	25	19	24
62.	There will be little rain when the points of the moon turn up	14	8	23
63.	A compass needle is a little magnet	17	9	22
64.	The electricity in a lightning discharge is of the same nature as the electricity we ordinarily use	37	47	30
65.	An electrical generator causes electricity to move rather than creating it	46	28	52
66.	Electricity does not flow readily through rubber	5	2	8
67.	Lightning never strikes twice in the same place	14	6	25
68.	One should never touch even a small piece of metal during a severe lightning storm	36	25	45
69.	Radio waves travel at the same speed as sound waves	26	14	37
70.	Radio waves need air to travel on	52	32	69
71.	Sound is always produced by vibration	3	2	2
72.	Sound requires a solid, liquid, or gas to travel through	30	10	44
73.	A telephone message is conveyed by a current of electricity	11	15	10
74.	Light travels instantaneously from the sun to the earth	27	9	52
75.	Light will travel through a vacuum	20	15	25
76.	Ether is a non-material substance which is found in all matter	30	36	22
77.	We see an object because particles travel from the eye to the object	17	9	26
78.	A shadow is a place from which light has been cut off	14	17	16
79.	The distant stars shine because they are hot like the sun	53	36	60
80.	The cat sees much better at night than do most other animals in the daytime	56	37	71
81.	On a very cloudy day no light from the sun reaches the surface of the earth	6	4	10
82.	The stars are not visible in the daytime because our side of the earth is turned away from the direction in which they are found	15	7	21
83.	The moon shines by light reflected from the sun	12	3	17
84.	The weather always clears up after a rainbow	29	17	36

INTERPRETATION OF RESULTS

An examination of the Science Inventory items will show that nearly all of them are stated positively so as to guard against errors made through mere mental confusion. Moreover, the fact that the items are concerned with ideas so simple in character that some understanding, belief, or bias will have been present in the minds of most subjects would tend to reduce the amount of "pure guessing." Nevertheless, a significant number of subjects may have engaged in pure guessing on some of the items; nor should the possibility of unintended interpretations of items be overlooked.

A. Chemistry

The possession, at an elementary level of comprehension, of some understanding of the present knowledge of the ultimate composition of matter might reasonably be expected of most beginning college students. However, 45 per cent of these subjects subscribe to a statement of the classical Greek conception of the elements as fire, air, earth, and water (1)[1]; while 48 per cent reject the statement that there are smaller units than atoms entering into the makeup of matter (2). Nor is it probable that the conception of a possible material unity to the universe has been developed by more than half of these subjects, if by that many, since 43 per cent deny the statement that matter is electrical in nature (3) and the same number reject the statement that the stars contain few, if any, elements not found on earth (4). Furthermore, the group of highest scholastic aptitude does not display much greater understanding of the composition of matter in their reactions to the items just considered than the group of lowest scholastic aptitude, as a glance at the table will show (1, 2, 3, 4). However, it should be noted that only 18 per cent of the entire group subscribe to the fallacious popular conception of atoms as "hard, solid, perfectly spherical bodies" (5).

Oxidation is one of the few chemical processes which, because of its relative simplicity and its significance in daily life, is found in nearly all, if not all, state courses of study in elementary science. It is also emphasized in most general science courses of study. It is consequently not surprising, although gratifying, to note that all but 9 per cent of the subjects apparently understand that ordinary burning involves a union of a substance with oxygen (6). However, many subjects' conception of oxidation does not go much beyond this understanding, since 35 per cent fail to recognize an explosion of coal dust as a case of rapid burning (7); while 40 per cent reject the statement that water is formed whenever most fuels burn (8) and 39 per cent apparently do not consider water to be an oxide (9). An even larger number—54 per cent—do not know that iron rust is heavier than the iron out of which it is formed (10).

[1] The numbers in parentheses are the numbers of the items in the table.

The understanding of oxidation by the group of highest scholastic aptitude is apparently not much greater than that of the group of lowest scholastic aptitude (6, 7, 8, 9, 10). However, the fallacious popular beliefs that air is a compound and that water becomes air when it evaporates are only half as prevalent in the group of highest scholastic aptitude as in the group of lowest scholastic aptitude (11, 12).

Belief in that crude form of "current creation" known as "spontaneous generation" is known to be widespread in the general population, although over half a century has elapsed since Pasteur and others succeeded in disproving every known supposed case of the phenomenon. This belief is found to exist extensively in this group of college students. It is impossible to determine from the data, however, just how many in total entertain the belief in some form or other, since some subscribe to the belief with reference to one kind of organism and some with reference to another (13, 14, 15, 16).

Verbal "knowledge" is, unfortunately, often present to a greater degree than important related functional knowledge; and this weakness exists with reference to these subjects' knowledge of at least one of the food nutrients—protein. While 99 per cent subscribe to a correct verbal statement of the functions of protein in the diet, 39 per cent reject the statement that the white portion of milk is protein in character, and 36 per cent reject the statement that the lean portion of meat and the white portion of milk serve about the same purposes in the diet (19, 20, 21). Attention should also be called specifically to the fact that about one-third of the entire group of students subscribe to statements of each of three widespread popular superstitions concerning foods and that these superstitions are much more prevalent in the group of lowest scholastic aptitude than in the group of highest scholastic aptitude (22, 23, 24).

Photosynthesis, representing as it does one of the most important chemical processes occurring on earth, might well be understood in an elementary way by most people. It is to be noted, however, that although all but 15 per cent of these subjects know that chlorophyll is essential to the growth of plants 43 per cent do not comprehend the important fact that the carbon dioxide of the atmosphere enters into the formation of plant material, but accept the statement that all of the material entering into the formation of a tree comes from the soil and the water (36, 39). It is gratifying to note, however, that all but 13 per cent grant the modest claim that man can manufacture some of the compounds which are made by plants and animals (38).

B. Physics

Beliefs resulting from uncritical, everyday observation sometimes operate against the acceptance of important laws and principles which have, of course, been established by critical experimentation. So the ordinary man has often noted that bodies stop soon after power is no longer applied, while his observation does

not, and could not, provide the reason for the cessation of motion in the opposing force of friction. He also notes that light, fluffy bodies normally fall at a slower rate than dense, heavy bodies, and his observation does not tell him that the difference in the behavior of the two kinds of bodies is due to the friction of the atmosphere. It is not surprising, consequently, that great weaknesses are revealed in these students' understanding of the general laws of motion and of the specific laws of falling bodies. Although a number of items are devoted to these topics, attention is specifically called to but a few. It is pointed out that 66 per cent accept the statement, whatever it means, that it is more natural for a body to stop itself than to set itself in motion (41); 37 per cent do not realize, apparently, that a uniform force tends to produce accelerated motion (42); and 46 per cent apparently think, as Aristotle did, that bodies fall with speeds proportionate to their masses (45). To be more fully effective physics teaching must take into account the influence of uncritical, everyday observation in the establishment of fallacious concepts and beliefs.

Correct statements of Archimedes' principle, as it applies both to floating and submerged bodies, are subscribed to by about 85 per cent of the subjects, and the success of the group of lowest scholastic aptitude is not much less than that of the group of highest scholastic aptitude (48, 49). Correct statements of the equality of liquid pressure, in all directions and in small bodies as compared to large bodies of liquid, are subscribed to by about 75 per cent of the total group; however, there are over three times as many in error in the group of lowest scholastic aptitude as in the group of highest scholastic aptitude (50, 51).

Illustration of the often-discovered educational situation in which knowledge of the bare fact is more extensive than understanding of the reasons behind the fact is provided by the results obtained on two statements concerning atmospheric pressure; for, while all but 20 per cent subscribe to the statement that the average pressure of the atmosphere at sea level is about 15 pounds against a square inch, 44 per cent reject the statement that the average weight of the atmosphere above a square inch at sea level is about 15 pounds (53, 54).

The old, classical belief that "nature abhors a vacuum" is held, apparently, by a larger number of the group of highest scholastic aptitude than of the group of lowest scholastic aptitude (55). A possible explanation is that the more intelligent students, because of their wider range of interests and intellectual activities, are more likely than the less intelligent to have been exposed to this fallacy.

The results on the items devoted to the topic of heat reveal one particularly startling weakness in the understanding by these subjects of the cause of the different seasons, since 79 per cent of the entire group of 325 subjects and even 64 per cent of the group of highest scholastic aptitude subscribe to the statement that the seasons are caused by the varying distance of the sun from the earth (59).

The comprehension of the fact that matter is electrical in nature would seem to be essential to an understanding of some of the simplest electrical phenomena; but, as has been previously pointed out, 43 per cent of the subjects fail to reveal a knowledge of this fact (3). It is consequently not surprising that 46 per cent apparently think that a generator actually creates electricity rather than setting in motion electrical particles which are already in existence (65).

An illustration of the important educational principle that the science teacher should be aware of the outside influences affecting the scientific thinking of his students is provided by the fact that 52 per cent accept the statement that radio waves need air to travel on (70). It is altogether probable that this misconception is engendered and reinforced by the repetition of the misleading phrase, "on the air," by radio announcers.

Support for the thesis that science teachers should go beyond the conventional and traditional boundaries of their subjects to promote the acquisition of important related valid ideas and to destroy those which are fallacious is provided by the results on two of the items concerned with the topic of light; for 53 per cent of the subjects do not know that the distant stars shine because they are hot like our sun, and 56 per cent subscribe to the fallacious popular belief in the superior night vision of the cat (79, 80). Study of the table will reveal many such items to science teachers.

SUMMARY AND CONCLUSIONS

Since this study is concerned chiefly with the prevalence of certain concepts and beliefs (both true and false) in the fields of chemistry and physics rather than with general accomplishment in these fields, few generalizations can be made concerning its findings. It may be said, however, that (1) the degree of prevalence of some correct concepts and beliefs and of some incorrect concepts and unfounded beliefs was ascertained with some degree of objectivity; (2) the degree of prevalence of true and false concepts and beliefs, even in the case of those closely related and equally significant, was found to vary widely; (3) the success of the group of highest scholastic aptitude on the individual concepts and beliefs was commonly greater, often significantly greater, than that of the group of lowest scholastic aptitude, although there were cases in which the success of the group of highest scholastic aptitude was no greater or even considerably less than that of the group of lowest scholastic aptitude; and (4) predictions based on scholastic aptitude of probable accomplishment on individual concepts and beliefs, or on individual items, would not have been highly reliable, judging by the great variation in results obtained from concept to concept and from related item to related item.

An Experiment in the Teaching of Inorganic and Analytical Chemistry[1]

WILLIAM E. MORRELL and NICHOLAS D. CHERONIS *Chicago City Colleges, Chicago, Illinois*

THE present experiment was undertaken as a result of round table discussions held by the Educational Group of the Chicago Section of the American Chemical Society during 1937–38.

The immediate objectives of the experiment were:

1. To eliminate from the freshman chemistry course repetition of material with which the student is already familiar, and to permit the unusually talented student to progress as rapidly as he can.
2. To test the progress that a good student can make when given the best possible conditions.

It is hoped that this study may lead ultimately to a third objective:

3. To determine what revisions might be desirable in the traditional freshman and sophomore college chemistry courses.

To insure objectivity an advisory committee was appointed, consisting of instructors from other institutions.[2] A general outline for the course was drawn up and submitted to the committee members for their criticisms. This original outline included the fields of inorganic chemistry and qualitative analysis, with emphasis on fundamental *principles*. Descriptive material was not emphasized, as it was assumed that the students would have considerable knowledge of descriptive chemistry from their high-school course. It was assumed that sufficient additional descriptive material would be obtained through assigned readings and as an incidental result of the study of chemical principles.

The first section of 32 students was enrolled in the class in September, 1938. The average grade of the students in their classes other than chemistry at Wright Junior College was exactly *B*.

The following textbooks were chosen because they emphasize principles rather than descriptive details: Hildebrand's "Principles of Chemistry," and Latimer and Hildebrand's "Reference Book of Inorganic Chemistry," and, as laboratory manual, Bray and Latimer's "Course in General Chemistry." For the work in qualitative analysis Hogness and Johnson's "Qualitative Analysis and Chemical Equilibrium" was also used.

The class had three discussion hours a week and six hours of laboratory. A difficult written quiz was given nearly every week, the time for the quiz being taken from either discussion or laboratory.

The first semester's work included kinetic molecular theory, atomic and molecular structure, the periodic system, ionic theory, chemical equilibrium, and oxidation-reduction. Incidental to these topics, hydrogen and Groups 0, I, II, V, VI, and VII of the periodic table were studied.[3]

The accompanying laboratory assignments (in Bray and Latimer) required individual work and more initiative on the part of the student than do the experiments in the majority of laboratory manuals. Analytical balances, burets, and pipets were used during the first semester. The experiments involved a comparatively large number of unknowns.[3]

In the second semester qualitative analysis and the study of the metals were taken up simultaneously. Each student analyzed eight unknowns by semimicro methods.[3]

At the end of the second semester a final examination[3] in general chemistry and qualitative analysis was constructed and administered by the advisory committee. The examination consisted of two parts: a three-hour written test and a laboratory examination. For the laboratory examination each student was given an unknown to analyze under the scrutiny of the committee. Following are pertinent excerpts from the report of the committee's conclusions:

> "All members of the committee are in agreement that . . . the students, as a whole, have mastered theory in a very satisfactory manner.
>
> "On the other hand, it should be pointed out that the descriptive material of the course was not emphasized. As a matter of fact it was thought by some members of the committee that the theory has been overemphasized in the course at the expense of the descriptive material.
>
> "The general opinion of the committee is that the laboratory technic of the class was fair. . . . For a selected group of students, it was thought that the laboratory technic should have been better. . . .
>
> "The committee unanimously agrees that this type of course should be continued next year."

The examination itself had included comparatively little descriptive material. It was suggested by the committee that the students' knowledge of descriptive matter be investigated further by giving the class the American Council on Education's Coöperative Chem-

[1] Part of the material in this article was presented before the Division of Chemical Education at the ninety-ninth meeting of the A. C. S., Cincinnati, Ohio, April 10, 1940.

[2] Professors W. C. Johnson and T. R. Hogness of the University of Chicago; L. F. Audrieth, J. C. Bailar, Jr., and B. S. Hopkins of the University of Illinois; H. T. Briscoe of the University of Indiana; R. K. Summerbell of Northwestern University; B. B. Freud of Illinois Institute of Technology; and P. M. Glasoe of St. Olaf College, of Minnesota.

[3] A detailed outline of the course is included in a monograph which may be obtained by addressing the writers at Wright City College, 3400 North Austin Avenue, Chicago, Illinois. The monograph also includes outlines of the laboratory work, copies of the final examinations given by the advisory committee, and reports by this committee.

istry Test for College Students. This test was given not only to these students but also to the beginning section of students in the experimental class. (These latter students had entered the course in February, and were nearing the end of their first semester of college chemistry. At the time of the test they had just finished general chemistry and were beginning qualitative analysis.)

The results of the two classes on the Coöperative Test, together with national norms,[4] supplied by the Coöperative Test Service, are listed in Table 1.

TABLE 1

COÖPERATIVE CHEMISTRY TEST FOR COLLEGE STUDENTS (Form 1939)

	I Information and Vocabulary	*II Problems and Equations*	*III Scientific Method*	*Total*
Maximum Possible Score	68	56	50	174
National Norms				
Percentile		Raw Scores:		
95	50	42	32	116
75	38	26	22	83
50	29	16	16	61
25	21	8	10	43
5	11	1	3	25
Students in Experimental Class				
Mean Scores:				
Advanced section	49.8	35.9	25.4	111.1
Beginning section	46.3	29.7	25.1	101.1
Percentiles:				
Advanced section	95	91	83	93
Beginning section	91	81	82	88

The national norms are for students who have had one year of college chemistry. The values in columns I, II, III, and Total are the scores at or below which fall the percentage of the cases indicated in the column at the extreme left of the table. For example, a student who made a total of 61 points on the test scored as high as or higher than 50 per cent of the students who took the examination.

The mean score of 111.1 for the advanced section of the experimental class signifies that it averaged as high as or higher than 93 per cent of the students in other colleges who took the test. Similarly the average score of 101.1 made by the beginning section after one semester of the experimental class indicates that it scored as high as or higher than 88 per cent of the students who took the test after a year of college chemistry.

The conclusion seems warranted that the emphasis on principles and theory had not been at the expense of "common chemistry." It had been assumed that the high-school background of these students could be relied on to a large extent for the factual and descriptive information which they should have, and that more would be picked up in the course without emphasis upon this phase of the subject. This assumption now seems fully justified.

A third section of students entered the class in September, 1939, and a fourth section began in February, 1940. These students were given the Coöperative Chemistry Test during their first week in the class. The third group, made up of the same 26 students who had taken the examination in September, 1939, was then retested, after one and one-half semesters of college chemistry in the experimental course. Similarly, the fourth group was retested after one semester in the experimental course. The results of these tests are shown in Table 2.

[4] J. CHEM. EDUC., **17**, 70 (1940).

TABLE 2

COÖPERATIVE CHEMISTRY TEST FOR COLLEGE STUDENTS (Form 1939)

	I Information and Vocabulary	*II Problems and Equations*	*III Scientific Method*	*Total*
Mean Scores				
Third Section:				
At beginning	35.6	18.2	20.9	74.7
After 1½ semesters	45.5	35.4	29.0	109.9
Fourth Section:				
At beginning	35.1	20.1	21.9	77.1
After 1 semester	45.3	35.5	29.3	110.1
Mean Score Percentiles*				
Third Section:				
At beginning	69	55	70	67
After 1½ semesters	89	90	91	92
Fourth Section:				
At beginning	68	60	74	69
After 1 semester	89	90	91	93
Means of Individual Percentiles†				
Third Section:				
At beginning	64	53	65	63
After 1½ semesters	84	83	88	88
Fourth Section:				
At beginning	63	57	68	65
After 1 semester	84	83	87	88

* Individual scores averaged, percentiles corresponding to these average scores then taken.

† Individual scores first converted to individual percentiles, these percentiles then averaged. (The data from which to make corresponding calculations for the first and second classes, whose results are given in Table 1, are not now available.)

This table shows that the mean scores of the third and fourth sections when they entered college corresponded to percentile ranks of 67 and 69, respectively, when compared with usual students who had had a year of college chemistry. The average score of the third section after one and one-half semesters in the experimental class corresponded to a percentile rank of 92, while that of the fourth section after one semester corresponded to a percentile of 93. Or calculated in another manner by first converting the individual scores to percentiles and then averaging these percentiles, the third section is seen to have improved from an average percentile rank of 63 to 88 in one and one-half semesters, while the fourth group raised itself from 65 to 88 in one semester.

It is possible that these newer sections had somewhat better high-school background in chemistry than had the earlier sections. The classes were, however, still no better than *B* average. In fact, in college courses other than this chemistry class the average grade of the students in the third and fourth sections was slightly below *B*.

The average percentile ranks of the newer sections compared with the original section, on the A. C. E. Psychological Test and in the high-school graduating classes, are shown in Table 3.

TABLE 3

AVERAGE PERCENTILE RANKS

Section	*1*	*2*	*3*	*4*
A. C. E. Psychological Test	74	72	81	82
High-School Graduating Class	82	82	83	82

As a result of more careful selection of students, improved organization of the course, and better budgeting of time, the fourth section progressed faster than did the first group. The material covered in the first semester by the first section was covered by the fourth section in three weeks less time.

Two months remained in the school year when the first experimental section completed qualitative analysis in April, 1939. At the suggestion of Dr. W. C. Pierce of the University of Chicago and with the approval of the advisory committee, the class then began quantitative analysis. The same unknowns[3] (seven in number) and the same methods of grading were used as are used by Professor Pierce in his beginning course at Chicago. Thirteen of the 26 students in the section finished the unknowns by the end of the semester.

The second section of students was similarly allowed to begin quantitative analysis in the latter part of its second semester and was less rushed than the preceding section had been.

The writers are confident that with the experience which had been gained the third group could have covered the work in quantitative analysis, as it was outlined, with little difficulty. However, with the approval of the committee, a more radical experiment was undertaken. The third group was given an integrated course in analysis during its second semester, a course which combined qualitative and quantitative analysis. Swift's "A System of Chemical Analysis" was used as text. The students first analyzed six quantitative unknowns, in order to become acquainted with basic quantitative technics. They then proceeded with the study of the metals and qualitative analysis, but certain constituents of the unknowns were obtained quantitatively as well as qualitatively.

The integrated combination of qualitative and quantitative analysis was not continued with the fourth group. The experience with the third group seemed to indicate that for proper development of such a course, a class even of superior students would need to devote more time than just the second semester of the freshman year. The procedure followed by the first and second experimental groups was returned to by the fourth group in its study of elementary analysis.

The third and fourth sections were examined (May 28, 1940) by the supervisory committee as the first section had been. The examinations[3] were made up and administered by the same subcommittee as before.

Lack of space here prevents discussion of the unique laboratory examinations[3] developed by the committee members and given by them to the students one week after the written tests. It is planned to make the examinations the subject of a future paper. Concerning these examinations the committee's report[3] stated: "Members of the committee had an opportunity to observe and question students throughout the three-hour examination period. It is the opinion of the committee that these students showed real aptitude in their attack on the problems and that their laboratory technic was excellent."

Following are pertinent excerpts from the summary:

> "The examining committee now believes that it is possible to formulate some rather definite recommendations on the basis of a two years' study of these experimental courses.
>
> "I. The experiment has proven a real success and should now be extended to a larger group of students. These students should, however, be chosen very carefully on the basis of a qualifying examination. . . .
>
> "II. The one-year course in general chemistry should make allowance for previous training in chemistry (high-school course) and emphasize primarily the teaching of principles of chemistry. . . .
>
> "III. The quantitative aspects of laboratory work should be emphasized as an application of, and correlated with the teaching of, principles of chemistry. This has been done in an admirable fashion but it is questionable if actual quantitative determinations, such as those comprising part of the laboratory work during the past year, should be continued. Part of this time should be devoted to further training in qualitative analysis. Such volumetric determinations as acidity, soda ash, and chloride content should still be retained, but more difficult determinations eliminated. . . ."

In keeping with the recommendations of the examining committee, the course is now being extended to larger groups of students. During the fall semester of 1940, of 215 students enrolled in beginning college chemistry, the upper 50 per cent—chosen on the basis of a placement test—were given the experimental course. The principal changes in the course were those made necessary by the larger number of students—lectures substituted for part of the discussion periods, and a somewhat slower pace. The two-semester course for the larger class includes inorganic chemistry and qualitative analysis. A number of the laboratory assignments are quantitative, but quantitative analysis as such is not included in the two-semester course. The course in general is being guided by the recommendations of the advisory committee.

In the meantime the earlier groups of students have taken organic chemistry. It has been found that they are better able to deal with theoretical aspects of organic chemistry than are students who have had the traditional course. However, on the memory work necessary in organic the experimental students find themselves at a disadvantage—unless they watch themselves—as they have lost the habit of memorizing, and prefer to rely on their knowledge of general principles.

In summary, to the conclusions and recommendations of the examining committee the writers add their own conclusions:

1. The traditional first two years' work in college chemistry can be profitably revised.

2. Good high-school chemistry can be relied on in large measure for background in descriptive material.

3. By emphasizing basic principles, considerable saving in time can be effected in elementary college chemistry and analysis. Not much time can be saved in qualitative laboratory, however, except by cutting down on the large number of unknowns sometimes given. Similarly, in quantitative laboratory, little

time can be saved on the unknowns themselves. Time spent on preliminaries can be cut down in an integrated course, though, by teaching the use of burets, pipets, and analytical balances in the general chemistry laboratory in place of experiments which are largely repetition of high-school work. In brief, general chemistry, qualitative analysis, and at least a certain amount of quantitative analysis can be adequately covered in two semesters.

4. Steps in the direction indicated can be taken by even the average class.

In conclusion, we wish to express to Dean William H. Conley of Wright City College and the committee members our sincere thanks for their interest in the experiment and for their wholehearted coöperation.

Classification of Analytical Samples

PHILIP J. ELVING *Purdue University, Lafayette, Indiana*

THE convenient classification and arrangement of the samples used in the various courses in quantitative analysis is a decided problem in many colleges and universities which possess large collections of analytical samples. In an attempt to systematize the arrangement of samples for quantitative analysis and to facilitate the keeping of records of these samples the classification shown in Table 1 was devised. The basis of the system of classification is the sample number, consisting in general of a letter and a three-digit number, *e. g.*, *D*643. The letter indicates the class to which the sample belongs, the first or hundredth digit shows the subclass or type of the sample and the last two digits give the number of the particular sample of the given type. The vertical column on the left of Table 1 gives the code letters indicating the classes of samples while the other vertical columns provide for ten types of samples in each class. The arrangement, in part, is based on the Periodic System.

The first eight classes of samples, *A* to *H*, provide for simple unknowns of the elements of the eight groups of the Periodic System, the elements in each group being indicated by the hundredth digit used. Thus, samples *B*300 to *B*399 are intended for the determination of calcium. Numbers 1 to 99 in each class are reserved for samples containing more than one member of the group, *e. g.*, *A*1 to *A*99 might be samples in which both sodium and potassium are to be determined. Numbers 100 to 599 are, in general, reserved for the elements of the principal subgroup, numbers 600 to 899 for elements of the other subgroup. Numbers 900 to 999 are reserved for samples in which the elements of the group are present or are to be determined as anions, *e. g.*, *G*600 to *G*699 would be samples of manganous salts while *G*900 to *G*999 could be samples for the determination of permanganate. Similarly, *F*200 to *F*299 could be samples for the determination of sulfide and *F*900 to *F*999 for the determination of sulfate. Usually, the samples in classes *A* to *H* will consist of simple salts diluted with inert material.

If more than a hundred samples of a given species are present, the numbering is continued by adding a fourth digit, going from, *e. g.*, *B*299 to *B*2001, *B*2002 up to *B*2999 if necessary. If several types of the same sample are present they may be differentiated by use of a fourth digit, *e. g.*, *H*1000 to *H*1099 for ferric salts and *H*1100 to *H*1199 for ferrous salts, or 1000 for chlorides, 1100 for nitrates, and 1200 for sulfates of the same metallic element.

The other sixteen classes of samples, *J* to *Z*, omitting *I* and *O*, provide for the numbering not only of the alloy and rock samples used in the ordinary elementary courses in quantitative analytical chemistry but also for the samples necessary for the usual courses in metallurgical analysis, technical analysis of all types, rock analysis, food analysis, agricultural analysis, quantitative organic analysis, spectroscopy, etc. Provision is also made for primary standards and various other special classes of substances. Ample provision is allowed for new samples and classes of samples. The classes or samples may be readily extended by using a double key letter, *i. e.*, *AA*, *AB*, etc.

In most cases there would be no real need to change the present numbers on samples in switching over to this or a similar classification of samples. It would be merely necessary to prefix a key letter and a key number to the present two- or three-digit number, aluminum salt sample No. 12 becoming *C*212. Where students may see into the sample storeroom the advantage of a sample being marked "*C*212" instead of "Aluminum 12" is evident.

Arrangement of the samples alphabetically and numerically in the storeroom greatly facilitates locating a sample of a given type and number, especially when a large number of samples containing only a few of each type are kept in the same location. Record-keeping is simplified since the sample records can be systematized on the same alphabetical-numerical basis.

TABLE 2

Key Letter	*Type of Sample*
A	Element
B	Simple binary salt or oxide
C	Simple salts, *e. g.*, metallic sulfates
D	Simple ores, *e. g.*, metallic oxides or sulfides
E	Complex ores, *e. g.*, phosphate rock
F	Alloys based on the given element
G	Commercial products where the given element is the principal constituent or the only constituent determined
H	Miscellaneous samples involving the given element

Note: The above samples may be diluted with suitable inert material.

An alternative system of classification of analytical samples is also based in part on the Periodic System, employing for the first eight classes of samples a two-letter, three-digit code, *e. g.*, *BD*372, where the first letter, *A* to *H*, denotes the group of the Periodic System while the second letter indicates the nature of the sample as shown in Table 2. The number locates the element in the group and the particular sample as in the scheme of classification previously described. Thus, *AF*600 to *AF*699 would be copper-base alloys while *EG*306 might be a sample of Paris Green in which only arsenic is to be determined. For classifying steels, fuels, agricultural products, etc., the original system is applicable.

TABLE 1

CLASSIFICATION OF ANALYTICAL SAMPLES

Sample Letter	Class	Sample Numbers 1–99	100–199	200–299	300–399
A	Group I		Li	Na	K
B	Group II		Be	Mg	Ca
C	Group III		B	Al	Ga
D	Group IV	$HCO_3 + CO_3$	C	Si	Ge
E	Group V		N	P	As
F	Group VI		O	S	Se
G	Group VII	Mixed halides	F	Cl	Br
H	Group VIII		Fe	Co	Ni
J	Standard samples, primary standards	Bureau of Standards samples	Bureau of Standards samples	Standards for spectroscopy	
K	Copper alloys	Brass and bronze	Brass	Bronze	Al-, phosphor-, and other bronzes
L	Light metal alloys	Be-base alloys	Mg-base alloys	Zn-base alloys	Cd-base alloys
M	"White metal" alloys	Sn-base alloys	Pb-base alloys	Sb-base alloys	Bi-base alloys
N	Resistant alloys	Ni-base alloys	Ni-Cr alloys	Co-base alloys	
P	Iron and steel	Iron	Plain carbon steel	Steel, containing Ni	Steel, containing Cu
Q	Alloy steels	Steel, containing W and Cr	Steel, containing W, Cr, and V	Steel, containing Mo	Steel, containing Mo, W, Cr, V
R	Ores, metallic content	Cu ores	Zn ores	Al ores	Pb ores
S	Ores, non-metallic content				
T	Silicates	"Silicates"	Feldspars	Orthoclases	
U	Carbonates, etc.	"Limestones"	Limestones	Dolomitic limestones	Dolomites
V	Fuels	Coal and coke	Petroleum crudes	Gasoline	Light burning oils
W	Agricultural and food products	Fertilizers	Feed stuffs	Insecticides	Soils
X	Miscellaneous technical products	Bleaching powders	Paints	Explosives	Textiles
Y	Samples for organic quantitative analysis	Metals and residues	Carbon-hydrogen	Nitrogen	Halogen
Z	Gases, waters, etc.	Single or "pure" gases	Flue gases	Fuel gases	Miscellaneous gases

Sample Letter	Sample Numbers 400–499	500–599	600–699	700–799	800–899	900–999
A	Rb	Cs	Cu	Ag	Au	H
B	Sr	Ba	Zn	Cd	Hg	
C	In	Tl	Sc and Y	Ce	Rare earths	B_2O_3 and borate
D	Sn	Pb	Ti	Zr and Hf	Th	CO_2
E	Sb	Bi	V	Cb	Ta	
F	Te	Po and U	Cr	Mo	W	SO_4
G	I		Mn	Ma	Re	
H	Ru	Rh	Pd	Os	Ir	Pt
J	Primary standards, acidimetry	Primary standards, alkalimetry	Primary standards, oxidimetry	Primary standards, reductimetry	Primary standards, precipitimetry and complex formation	
K	Cu-Ni alloys	Cu-Ni-Zn alloys	Cu-Mn alloys	Cu-Si alloys		
L	Al-base alloys					
M	Solders	Babbitts	Fusible alloys	Pyrophoric alloys		
N					Ag, Au, and Pt alloys	
P	Steel, containing Si, Mn	Steel, containing Cr or V	Steel, containing Cr and Ni	Steel, containing Cr and V	Steel, containing "*X*"	Steel, containing "*X*" and "*Y*"
Q	Steel containing Mo and "*X*"				Ferro-alloys	Alloying metals
R		Cr ores	Mn ores	Ni ores	Fe ores	
S	Phosphate rock	Sulfide ores				
T	Slags	Cements	Glass	Silica refractories	Glazes, enamels, etc.	
U	Magnesites	Burned limestone, lime, magnesia, etc.	Boiler scale	Soda ash		"Insoluble-in-direct"
V	Lubricating oils	Fuel and road oils	Greases	Waxes	Bituminous products	
W	Foods	Dairy products	Meat products	Carbohydrates and starches	Oils and fats	Beverages
X	Rubber, paper, etc.			Organic chemicals	Metallic salts	Non-metallic chemicals
Y	Sulfur	Molecular weight	Neutralization equivalent	Functional groups		
Z	Potable water	Boiler feed water	Sewage and effluents	Miscellaneous waters	"Moisture determination"	Specific gravity, density

The Gram Electron

L. F. AUDRIETH and M. J. COPLEY *University of Illinois, Urbana, Illinois*

THE terms *gram atom* and *gram molecule* are introduced early in the course in general chemistry to emphasize above all the quantitative meaning of chemical symbols, formulas, and equations. The term *gram ion* may also be introduced later, possibly in connection with a discussion of partial equations, such as (1) to (4), which are used to emphasize the meaning of electrovalence and to demonstrate the ionic nature of compounds formed when the indicated electron transfers (gains or losses) occur.

$$K \longrightarrow K^+ + \epsilon^- \quad (1)$$

$$Mg \longrightarrow Mg^{++} + 2\epsilon^- \quad (2)$$

$$[F] + \epsilon^- \longrightarrow F^- \quad (3)$$

$$[O] + 2\epsilon^- \longrightarrow O^{--} \quad (4)$$

Such half equations may also be given a quantitative meaning, even though the respective processes are not capable of realization by themselves under ordinary circumstances. Thus, one gram atom of potassium is converted into a gram ion of potassium in a process involving the loss of approximately 6.06×10^{23} electrons. It is proposed that this number and quantity (mass in grams) of electrons be designated the *gram electron*.

The question may properly be raised: Is the introduction of this new term justifiable? Can it be used to simplify the presentation of concepts whose relationships are not ordinarily evident to the beginning student? Actual and successful teaching experience, particularly with students who have had a high-school course in chemistry, has prompted us to submit this idea through the JOURNAL OF CHEMICAL EDUCATION to other teachers of chemistry.

WHAT IS THE GRAM ELECTRON?

The use of the term gram electron emphasizes the transfer of ponderable matter in all oxidation-reduction reactions, whether chemical or electrochemical. The gram electron represents approximately 0.00055 gram of matter—a not insignificant quantity. This figure is calculable from the mass of the electron (in the rest state), 9×10^{-28} g., multiplied by Avogadro's number, 6.06×10^{23}, or from the relative (atomic) weight of the electron compared with that of hydrogen. Just as the gram atom and the gram molecule represent definite quantities (masses) and numbers (6.06×10^{23}) of atoms and molecules, just so does the gram electron represent a related quantity (mass) and the same number of electrons.

THE GRAM ELECTRON AND THE FARADAY OF ELECTRICITY

The relationship to and the identity of the gram electron with the faraday can best be demonstrated by a consideration of the accepted definitions of a number of electrical units. Thus, the ampere is defined as that current which in one second deposits 0.001118 gram of silver from a silver nitrate solution. The coulomb is that quantity of electricity transferred by one ampere in one second. The coulomb may therefore also be defined as that amount of electricity required to deposit 0.001118 gram of silver from a silver nitrate solution. The number of coulombs which are needed to deposit 107.88 grams of silver (one gram atom) is equal to $\frac{107.88}{0.001118}$ or 96,494 coulombs. This number of coulombs has been found by experience to be the smallest amount of electricity which will produce an electrochemical change in a gram atom or gram molecule of any chemical species. It has become convenient to refer to this quantity of electricity as the "faraday," but it can just as properly be called a "gram atom of electricity," since the atom of electricity is the electron, or more briefly the "gram electron" as we propose. The gram electron and the faraday are therefore one and the same unit and quantity.

Many students experience considerable difficulty in visualizing what is meant by "quantity" of electricity. In emphasizing the fact that it is matter which is transferred this uncertainty disappears. If equations such as (5, 6) are interpreted as we propose, the meaning of quantity of electricity becomes perfectly obvious.

$$\underset{\text{one gram ion}}{Ag^+} + \underset{\text{one gram electron}}{\epsilon^-} \longrightarrow \underset{\text{one gram atom (107.88 g.)}}{Ag} \quad (5)$$

$$\underset{\text{one-half gram ion}}{{}^1/_2Cu^{++}} + \underset{\text{one gram electron}}{\epsilon^-} \longrightarrow \underset{\text{one-half gram atom (31.79 g.)}}{{}^1/_2\,Cu} \quad (6)$$

The gram electron concept is furthermore useful in interpreting reactions which occur at the poles of electrolytic and voltaic cells. Electronic equations such as those given above are not only qualitative, but also quantitative expressions in that they relate electrochemical change to the quantity of electricity and indicate the number of gram electrons involved per gram atom, gram ion, or gram molecule.

THE GRAM EQUIVALENT

The equivalent weight, or the gram equivalent, is that quantity of substance whose formation or subsequent reaction involves the loss, gain, transfer, or sharing of one gram electron.[1] This is certainly a much more fundamental definition than "that quantity of matter which reacts with, contains, or displaces one gram of hydrogen, or eight grams of oxygen, or

[1] It has been pointed out to the authors that this definition is a reversal of the logical derivation of the gram electron concept. The gram electron should really be defined in terms of the gram equivalent weight, since the faraday was first characterized as that quantity of electricity which will discharge 107.88 grams (one gram equivalent) of silver from solution.

35.46 grams of chlorine"—or any other convenient equivalent quantity of any of the other 92 elements. By using the gram electron definition it makes little difference whether the equivalent weight of an oxidizing or reducing agent, of a molecule or of an atom is desired. The equivalent weight of potassium sulfate, an ionic compound, is a weight in grams equal to the molecular weight divided by two, since the transfer of two gram electrons from two gram atoms of potassium was involved in the formation of one formula weight of the compound. The equivalent weight of hydrogen chloride is equal to the molecular weight taken in grams, since each atomic species has contributed one gram electron in the formation of one mol of the covalent compound. The equivalent weight of zinc as a reducing agent is the atomic weight divided by two, since one gram atom of zinc may lose two gram electrons (equation 7).

$$Zn \longrightarrow Zn^{++} + 2\epsilon^- \qquad (7)$$

The equivalent weight of permanganate as an oxidizing agent in acid solution is equal to the molecular weight divided by five since one mol of permanganate takes up five gram electrons as indicated by the following half reaction.

$$MnO_4^- + 8H^+ + 5\epsilon^- \longrightarrow Mn^{++} + 4H_2O \qquad (8)$$

The use of this new term has justified itself in our own teaching experience. We have found it especially helpful where students already have had enough of a background in chemistry, so that the semester's work can be started with a thorough discussion of atomic structure. The gram electron is then introduced along with a discussion of the chemical significance of our modern concepts of atomic structure. Formation of electrovalent compounds by transfer of electrons is explained by use of partial (electronic) equations. This gives an opportunity to bring in oxidation and reduction, both chemical and electrochemical, from a quantitative point of view by using the gram electron concept. The redefinition of equivalent weight follows in logical sequence.

The authors take pleasure in acknowledging their indebtedness to Professor T. E. Phipps of the University of Illinois and Professor P. W. Selwood of Northwestern University for their valuable suggestions.

A USEFUL CHART FOR QUALITATIVE ANALYSIS

L. P. BIEFELD

Purdue University, Lafayette, Indiana

CHARTS displaying flow diagrams, curves, tables, and apparatus sketches are valuable aids to lectures and class discussions in qualitative chemical analysis. Satisfactory displays are drawn on 2′ × 3′ white Bristol board with black and red India inks.

The drawings are not substitutes for lantern slides and demonstration experiments but are used in conjunction with them. Charts have advantages over slides. No special projection or light control equipment is required. A series of charts can be exhibited and discussed simultaneously. Being able to refer rapidly back and forth from one drawing to another facilitates comparisons, contrasts, and explanations of visual material presented.

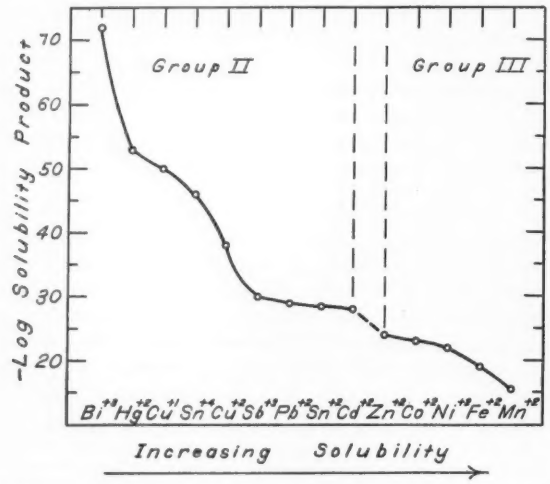


FIGURE 1.—SOLUBILITY PRODUCT VALUES OF THE SULFIDES OF GROUPS II AND III

Figure 1 shows a chart helpful for discussion of solubility product, solubility, and controlled precipitation as applied to separation of the acid sulfide from the alkaline sulfide group. The curve is obtained by plotting the negative logarithms of the solubility products of the sulfides against the cations involved. The cations are in order of increasing solubility of their respective sulfides and are separated by an arbitrary, uniform distance.

The solubility product values were procured from the class textbook (*1*). The values for the alpha modifications of cobalt and nickel sulfides were chosen because these forms presumably predominate when Group III is precipitated. The solubility products of ferric and arsenious sulfides are omitted because the values in the literature are not closely correlated to the actual conditions of precipitation (*2*).

LITERATURE CITED

(1) MIDDLETON AND WILLARD, "Semimicro qualitative analysis," Prentice Hall, Inc., New York City, **1938**, p. 111.
(2) *Ibid.*, p. 107.

Demonstrations of Gas Explosions

W. E. THRUN and J. M. LIEN *Valparaiso University, Valparaiso, Indiana*

EXPLOSIONS of hydrogen and other gases with oxygen or air have been demonstrated using rubber balloons (*1*, *5*), tubes (*1*, *6*), soap bubbles (*7*), and cardboard containers (*4*). Here are presented directions for showing differences in violence of explosive mixtures in bubbles, balloons, bottles, and other containers. The method of proportioning gases depends upon the fact that the volume of a sphere is proportional to the diameter. It has the advantage that the change to different gas ratios can be made quickly.

EXPLOSIONS IN BUBBLES

For blowing bubbles the triethanolamine oleate solution prepared as recommended by Cook (*3*) is the best we have used. His Castile soap solution is satisfactory. Three per cent solutions of the newer wetting agents with 10 per cent to 15 per cent glycerol give fairly tough bubbles and a combination of high and low foaming power wetting agents with glycerol may be used. A typical formula is: 10 g. Aerosol OT aqueous 10% (American Cyanamid and Chemical Company), 2 g. Tensol (Synthetic Chemicals, Inc.), 10 g. glycerol, and 80 g. water.

The explosive gas-oxygen mixture may be passed from a small rubber balloon (for preparation see below) into the pipe. *Be sure to free* bubble or *disconnect* the pipe from the supply *before* igniting in order to prevent an explosion in the balloon.

A method of preparing bubbles that can be safely exploded on the pipe is as follows: Inflate one small balloon with the gas from a cylinder or closed-system generator and another with oxygen. Connect them with rubber tubing to a T-tube, using pinch clamps or stopcocks to shut off the gases. Connect a clay or glass pipe with enough rubber tubing so that the combined volume is 10 ml. to the third opening of the T-tube. (Since oxygen sweeps the gas from the tube and pipe into the bubble the 10 ml. have been taken into account in calculating the diameters required to give theoretical mixtures.) Blow the bubbles with the gas to 2.5-cm. diameter and then pass in oxygen. For the hydrogen explosion the final diameter should be 4.20 cm., for hydrogen sulfide 4.97 cm., for methane 5.27 cm., for acetylene 5.55 cm., for ethylene 5.81 cm., and for ethane 6.05 cm. Measure these diameters with a rule or calipers to 0.2 cm. to insure loud reports. Blow bubbles upside down if they become heavy. Gas-air explosions are mild. For hydrogen and acetylene the final diameters are 5.50 cm. and 8.63 cm.

EXPLOSIONS IN BALLOONS

In connection with the use of balloons the following should be noted: (*a*) In order to force a gas from one balloon into another it is often necessary to stretch the receiving balloon by hand in order to overcome the initial resistance. (*b*) It is easier to force gas from a less to a more highly inflated balloon. (*c*) Spherical balloons of different uninflated sizes hold very nearly the same volume when inflated to the same diameter. This is borne out by measurements of volumes at atmospheric pressure by displacement of water. Then the pressure of the gas within the balloons must be practically the same at different degrees of inflation. A water manometer was connected through a rubber stopper to inflated balloons of various makes and sizes. Diameters and pressure readings were taken as they were deflated and in each case the pressure was greatest at moderate inflation, then less at higher inflation, and greater again at still higher inflation. The difference in pressure in the same or in different balloons is not likely to exceed 15 mm. mercury, causing an error of less than two per cent. (*d*) Acetylene escapes from balloons faster than hydrogen. For example: Two sets of balloons of the same size and make were inflated with hydrogen and acetylene to 12-cm. diameters and kept in the same room. At intervals diameters were measured and from them the loss in volume was calculated. The hydrogen loss was 6 per cent and the acetylene loss 30 per cent per hour.

Use small rubber balloons of 5- to 8-cm. diameter when uninflated and shaped to be spherical when inflated. Insert a short glass tube into a No. 2 rubber stopper until one end of it is flush with the larger end of the stopper. Force the larger end of the stopper into the neck of the balloon until the rubber overlaps the smaller end. Thus the balloon will have less of the neck inflated and take on a more nearly spherical shape. Attach rubber tubing with a pinch clamp to the glass tube. Set a large pair of calipers at proper distance to serve as guide for inflation. Work with two balloons, filling one to a diameter of 10 cm. with the combustible gas (H_2, H_2S, CH_4, etc.), and forcing this into another balloon filled to the necessary diameter with oxygen or air. For hydrogen, the oxygen balloon should have a diameter of 7.95 cm. (or air 13.44 cm.); for hydrogen sulfide diameters are 11.45 and 15.58; for methane 12.60 and 17.7; for acetylene 13.56 and 22.96; for ethylene 14.42 and 24.38; for ethane 15.18 and 25.66. Acetylene gives satisfactory explosions with air but it must be exploded without delay. The filled balloon is tied up with a soft cord, attached to a meter stick, and exploded by dangling it in the flame of a candle or burner in an open space. *Warning:* A theoretical acetylene-oxygen mixture should not be exploded in a small room.

EXPLOSIONS IN BOTTLES

To show that explosions in confined spaces are more violent than in bubbles or balloons, wide-mouthed bottles are used. Push two iron rods or nails through

two small holes of a tightly fitting stopper. Connect the ends which will project into the bottles with a little steel wool or a fine wire. When the bottle filled with the explosive mixture is closed tightly and an extension cord is connected to the iron rods or nails, place a heavy towel or cone-shaped wire screen over the bottle. If the cord is connected to an electric outlet, the wire will become hot and set off the explosion. Rubber stoppers are blown out and cork stoppers are broken up. Six-ounce bottles are used for gas-oxygen mixtures and liter bottles for gas-air explosions, the gases being prepared in balloons and forced into the bottles by displacement of water. Gasoline-air or ether-air explosions are successful if 0.14 ml. of gasoline or 0.16 ml. of ether is placed in the liter bottles.

Gas-air and gas-oxygen explosions can also be performed with the gas generated in a bottle. Drop the weight of zinc, ferrous sulfide, calcium carbide, or aluminum carbide calculated to give a theoretical mixture with the volume of air or oxygen directly into the bottles containing a small volume of a solution of acid or water, and stopper quickly. For comparing oxygen and air mixtures, using the same volume of gas, bottles of the correct volume usually cannot be found but the bottle volumes can be adjusted with pebbles or water. If hydrogen is generated in the bottle containing the theoretical amount of oxygen the pressure (three atmospheres) is so great that the stopper is usually blown out. To avoid this, fill a larger bottle with dilute acid until the desired volume for oxygen remains. A glass tube reaching through a hole in the stopper nearly to the bottom allows the acid solution to be displaced as hydrogen is generated. When the desired volume of liquid has been displaced the outlet is closed off.

ADDITIONAL DEMONSTRATIONS

Most of the following experiments can be done with the explosive mixture obtained by forcing one balloon of ethylene of 10-cm. diameter into an oxygen balloon of 14.4-cm. diameter. Explosive mixtures of other gases may be used instead of ethylene.

Allow an explosive mixture to escape from a balloon under water at such a rate that bubbles emerge in rapid succession. The flame of a burner is played upon the emerging bubbles. The series of explosions resulting suggests a machine gun. Prevent accidents by releasing the bubbles at such a depth that there is no possibility for the explosion to be carried back to the balloon.

Pass the explosive mixture from a balloon into an aqueous solution of soap or a wetting agent of high foaming power contained in a large mortar or pan. As soon as the foam is of maximum height remove the balloon and explode with a long metal gas taper. The explosion is much more violent than one in a balloon of the same volume.

Pass an electrolytic mixture of chlorine and hydrogen directly from a cell into a small dish of wetting agent solution. Remove and explode the foam. Direct winter sunlight through glass windows does not explode bubbles of the mixture.

To show that explosive oxygen-gas mixtures can travel in small-bore tubing with shattering results, connect a one-foot length of 4- to 5-mm. bore glass tubing loosely with rubber tubing to another length, 4 to 6 feet long. Provide the ends with rubber tubing and pinch clamps. Run the explosive mixture through the tube and quickly close the ends. Place a towel over the longer piece of tubing. Pull the shorter piece of tubing away from the longer and without delay explode with a burner flame. The explosion sounds like the report of a rifle. Under the towel a part of the tube will be found shattered. Explosions of mixtures such as electrolytic hydrogen-oxygen may travel in tubes down to 0.5-mm. bore. For explosions in rubber tubing of 5- to 7-mm. bore no towel is needed. The tube seems to undergo a convulsion.

Rubber balloons in place of the Meker burner (*6*) used to introduce gas-air mixtures into tubes for flame speed demonstrations allow the composition to be known more definitely. Connect three lengths of 22- to 26-mm. bore tubing with tape. Provide stoppers for both ends. Insert a one-holed stopper into one end. Pass gas-air mixture quickly from a balloon through the tube until an assistant at the other end smells the gas. Shut the gas off and without delay stopper both ends and take the balloon away. Open one end and start the flame with a burner. For the preparation of lean or rich mixtures no great variations from the theoretical composition are advisable. A theoretical ethylene-air mixture blue flame in a 22-mm. bore tube traveled 15 feet in three seconds.

For an explosion due to a combination of "defective wiring" and a gas "leak" connect with fine wire two nails which pass through a wall of a rigid and tight cardboard or wooden box or "house" of 4- to 20-liters volume. Cut a hole for a door, using tape for hinges. Place a closely fitting board with a gable roof on the house. Now pass through a hole in the wall the volume of combustible gas required to give a theoretical gas-air mixture. Stopper the hole and connect the "wiring" to the A.C. circuit. The roof and ceiling are blown off and the door blown open.

For gasoline explosions warm one or two spot plates and set them upon the floor of the house. Pipet about 0.13 ml. of gasoline for every liter of volume through the hole upon the plates. Explode after about one or two minutes.

For a "gas tank" explosion cut a round hole in the side of an ether can by center punching and boring with a countersink. Wire the can to a board. Pipet about 0.13 ml. gasoline per liter volume into the can and stopper the hole. After allowing the gasoline to diffuse for a time remove the stopper and play the flame of a gas taper around the opening. There is no loud explosion but a rush of the gases out through the hole.

LITERATURE CITED

(1) ARTHUR, "Lecture demonstrations in general chemistry," 1st ed., McGraw-Hill Book Co., Inc., New York City, 1939, pp. 99, 100.
(2) CLARK, J. CHEM. EDUC., 17, 472 (1940).
(3) COOK, *ibid.*, 15, 161 (1939).
(4) FINE, *ibid.*, 8, 929 (1931).
(5) LISK, *ibid.*, 7, 2919 (1930).
(6) THRUN, *ibid.*, 13, 136 (1936).
(7) ULLMAN AND HAZLEHURST, *ibid.*, 11, 113 (1934).

SOLUBILITY GENERALIZATIONS[1]

CHARLES J. HEIMERZHEIM *Brooklyn College of Pharmacy, Brooklyn, New York*

UNLESS otherwise specified, the statements apply to the "normal" salts of the following ions which are not decomposed by water: Na, K, NH_4, Ca, Sr, Ba, Mg, Al, Mn, Zn, Cr, Fe, Cd, Co, Ni, Sn, Pb, Cu, Bi, As, Sb, Hg, Ag. "Soluble" has been used to indicate a solubility greater than "5 grams of salt in 100 ml. of water" at room temperature.

SALTS OF MONOBASIC ACIDS

Chlorides, Bromides, Iodides: Soluble, except Ag, Hg (ous), Pb, Cu (ous); and iodides of Hg (ic), Bi, Sn

Hypochlorites, Chlorites, Chlorates, Perchlorates: Generally soluble, except Ag and Pb chlorites; $KClO_4$

Bromates: Similar to bromides, except Ba, which is insoluble

Iodates: Insoluble, except Na, K, NH_4, Mg

Fluorides: Generally insoluble, except Na, K, NH_4, Ag, Hg (ous), Al, Sb, Sn, Bi

Nitrites and Nitrates: Soluble, except Ag nitrite

Permanganates: Soluble, except Ag

Cyanides: Generally insoluble, except Na, K, NH_4, Ba, Sr, Ca, Hg (ic)

Thiocyanates: Generally insoluble, except Na, K, NH_4, Ba, Sr, Ca, Fe, Zn, Mn (ous)

Hypophosphites: Generally soluble, except Fe. (*Note:* This acid is not monobasic, but it acts as one, forming "primary" salts almost exclusively.)

Acetates and Formates: Generally soluble, except Ag, Hg (ous), Bi, Cr (ous) acetates and Hg (ous), Pb, Fe (ous) formates

SALTS OF POLYBASIC ACIDS

Sulfides and Sulfites: Insoluble, except Na, K, NH_4; $FeSO_3$

Sulfates: Generally soluble, except Ag, Hg (ic and ous), Pb, Ba, Sr, Ca, Sb (ous), Fe (anhydrous)

Thiosulfates: Soluble, except Ag, Pb, Ba

Ferricyanides: Insoluble, except Na, K, NH_4, Ca

Ferrocyanides: Generally insoluble, except Na, K, NH_4, Ca, Sr, Mg

Phosphates (Ortho-, Meta-, and Pyro-; Normal and Secondary): Insoluble, except Na, K, NH_4

Arsenates and Arsenites: Insoluble, except Na, K, NH_4 arsenate, Mg arsenite

Chromates: Generally insoluble, except Na, K, NH_4, Ca, Mg, Sn (ic)

Borates, Silicates, Molybdates, Tungstates: Insoluble, except Na, K, NH_4; $BaSiO_3$ and $Mg_3(BO_3)_2$; some alkaline borates only slightly soluble

Oxalates, Tartrates, Citrates: Insoluble, except Na, K, NH_4; antimonous tartrate and ferric oxalate

Carbonates, Oxides, Hydroxides: Insoluble, except Na, K, NH_4; $Ba(OH)_2$

GENERAL OBSERVATIONS

1. Practically all Na, K, and NH_4 salts are soluble.
2. Most of the commonly used salts of monobasic acids are soluble. (Fluorides, cyanides, thiocyanates, and iodates are exceptions.)
3. Most of the salts of polybasic acids are insoluble (except those of Na, K, and NH_4).
4. "Activity series" relationships and those in the periodic families of metals are apparent among the "exceptions."
5. Nearly all Ag salts are insoluble, except the chlorate, perchlorate, fluoride, and nitrate.

SPECIAL NOTES

1. Hydrolysis has an important effect on solubilities, often causing partial precipitation. (Salts composed of anions of "weak" acids, with tri- and tetravalent cations, are highly hydrolyzed.)
2. "Basic" or "sub" salts are practically all insoluble. (Among the few soluble ones are the basic nitrate and acetates of Pb; and the oxychlorides of Sn.)
3. Nearly all As, Sb, and Bi salts are either insoluble or decomposed in water, except Sb (ous) tartrate, SbF_3, SbF_5, $SbCl_3$, As_2O_5, AsF_5.

[1] Material collected from:

HODGMAN AND HOLMES, Editors, "Handbook of chemistry and physics," 24th ed., Chemical Rubber Company, Cleveland, Ohio, 1940. (Only salts listed in this handbook are considered.)

TREADWELL, "Analytical chemistry," based on German text, translated and revised by HALL; Vol. I, "Qualitative analysis," 9th English ed., John Wiley and Sons, Inc., New York City, 1937, 630 pp.

MCALPINE AND SOULE, "Fundamentals of qualitative analysis," D. Van Nostrand Co., Inc., New York City, 1936, 325 pp.; "Introduction to qualitative chemical analysis," D. Van Nostrand Co., Inc., New York City, 1939, 118 pp.

A Simple Demonstrating Ultramicroscope

HAROLD J. ABRAHAMS *Central High School, Philadelphia, Pennsylvania*
WILLIAM BLITZSTEIN *University of Pennsylvania, Philadelphia, Pennsylvania*

THE colloidal state of matter is of both scientific and industrial importance. Yet most teachers of chemistry are unable to have their students actually see colloidal particles in motion because ultramicroscopes are so expensive. It is hardly necessary to advance an argument in favor of letting students observe the Brownian movement with their own eyes rather than giving them a word picture of this phenomenon. A view of a colloid through an ultramicroscope will produce a stimulating and lasting impression upon students' minds, while a mere description will hardly have this effect. These considerations led us to devise the ultramicroscope herein described.

Essentially, an ultramicroscope consists of two parts: a microscope and a dark field illuminator. The microscope in this case is an ordinary instrument with magnification of 150× or greater. The illuminator is simple, compact, and particularly adapted for demonstration purposes.

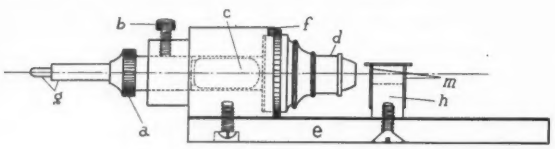


FIGURE 1.—SCHEMATIC DIAGRAM OF ULTRAMICROSCOPE ILLUMINATOR

a. Lamp socket
b. Lamp socket set screw
c. Lamp bulb
d. Condenser
e. Base of illuminator
f. Set screw for condenser
g. Contacts for electricity
h. Cell
m. Glass windows

As it is the illuminator which embodies most of the interesting features, it will be described in detail (Figure 1). Its function is to produce within the colloid an intense beam of light which comes to a focus which is not a point but a straight line (an effect similar to the Tyndall cone). The optical axis of the viewing microscope should be perpendicular to this line and the instrument should be focused on it (Figure 2). The effect of this type of lighting is to illuminate only those colloidal particles in the very thin section of the colloid which are in focus. Thus those particles out of focus receive no light and do not flood the field with diffuse illumination. Then each colloidal particle in the beam diffracts the light falling on it into the microscope and appears as a tiny luminous point on an essentially dark field.

The optical design of the illuminator is novel in that it consists of only two parts: a lamp and a lens. To produce the bright line most ultramicroscopes illuminate

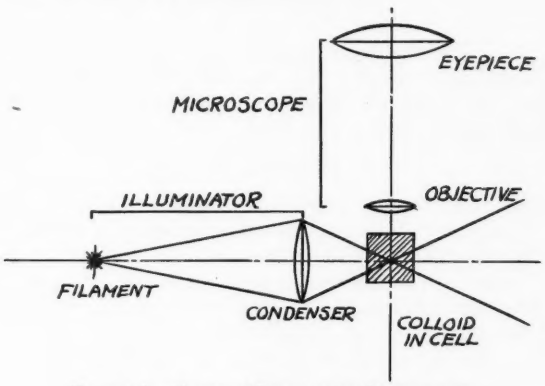


FIGURE 2.—OPTICAL DIAGRAM OF INSTRUMENT

a slit by an arc light and a condenser. A lens forms a brilliant real image of this slit in the colloid. We have replaced the illuminated slit by its optical equivalent, a bright filament which is in the form of a straight line. (The lamps used are those which are made for the ophthalmic instrument known as the "streak-retinoscope.") A low-power microscope objective is used to produce an image of the filament in the colloid.

The cell or container for the specimen must be transparent on at least three sides (Figure 3). It must allow the beam from the illuminator to enter and leave the colloid with a minimum of diffusion and internal reflection and it must allow the observer to look down on the beam. In our instrument, a semicylindrical slot was

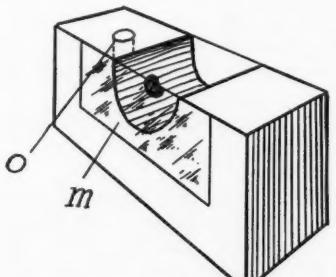


FIGURE 3.—PERSPECTIVE VIEW OF CELL

m. Glass windows
o. Intake for filling cell

milled in a block of brass. The ends of the slot were covered by microscope cover glasses sealed to the block

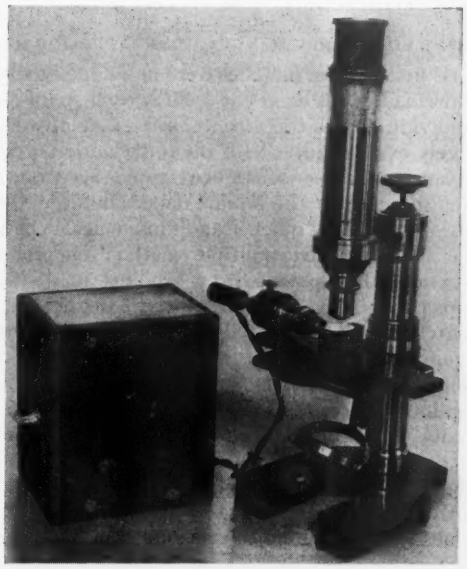
with paraffin. (Microscope cover glasses were used because they are so thin they have a negligible effect on the optical system.) This made two windows for the beam to enter and leave. A cover glass laid on top of the cell closed in the colloid and eliminated the vibration of the free surface of the liquid. The whole cell was lacquered dull black to reduce diffusion of light.

The instrument is small enough to rest on the stage of any ordinary microscope and can be easily moved about, which facilitates adjustment for observation. The electricity for the lamp (2.5 volts) is supplied by a small step-down transformer from the 110-volt A.C. mains.

Colloidal particles in both liquids and gases may be examined. At the low magnifications used, the visibility of the Brownian movement is slight in liquids—for example, various gold sols[1] and India ink; however, we have observed it very distinctly and easily at 150×. As viewed in our instrument, the particles in these gold sols appeared very similar to those pictured in 3 and 4 of Plate 1 in Zsigmondy's book on colloids.[2] In gases (smoke, etc.) the motion is quite violent and astonishing, offering a striking demonstration of the kinetic theory. In both cases it is very easy to see the particles themselves.

[1] "Gutbier's Gold," "Gold by Pyrocatechin," "Von Weimarn's Gold," "Zsigmondy's Gold." These were prepared from directions given in HOLMES, "Colloid chemistry," John Wiley and Sons, New York City, 1934.

[2] ZSIGMONDY-ALEXANDER, "Colloids and the ultramicroscope," John Wiley and Sons, New York City, 1909.



Courtesy of Mark H. Hagmann, Mathematics Department, Gratz High School, Philadelphia

DEMONSTRATING ULTRAMICROSCOPE

AN HISTORIC PIECE OF APPARATUS

JOSEPH Priestley's famous burning glass is here shown as it was moved from the Tome Scientific Building to its place of honor in a fireproof vault in the new Bosler library as a part of the 168th commencement exercises at Dickinson College recently.

For over a century Dickinson College has treasured certain pieces of apparatus once owned by Dr. Joseph Priestley, discoverer of oxygen, including the burning glass through which oxygen was probably first liberated.

Priestley came to America in 1794, when Dickinson was first feeling the inspiration given by Dr. Charles Nisbet, who came to it in 1785, from Edinburgh, as one of the most learned men of his time. Priestley's friend, Thomas Cooper, who became Professor of Chemistry at Dickinson in 1811, had lived with him in Northumberland, Pennsylvania, and it was through him that, in 1812, the precious historic apparatus referred to became the possession of Dickinson College. Particularly notable is the burning glass with its two mounted lenses, one sixteen inches and the other seven inches in diameter, set sixteen inches apart. Using this on August 1, 1774, Priestley focused the sun's rays upon a tube containing "mercurius calcinatus *per se*" (HgO) and thus obtained a new gas "in which a candle burned with a remarkably vigorous flame," and which supported life. He called this gas "dephlogisticated air." Some years later the French chemist Lavoisier named it oxygen.

PRIESTLEY'S BURNING GLASS AT THE DICKINSON COLLEGE COMMENCEMENT

A History of the Match Industry[1]

M. F. CRASS, JR. *Washington, D. C.*

PART VIII—EARLY MANUFACTURING PROCEDURE

THE first friction matches were individually dipped and made with hand-cut splints, each splint being individually whittled from a soft grade of wood of relatively even grain. With the introduction of white phosphorus and the consequent rapid acceptance of friction matches by the public, it became necessary to produce splints on a quicker and more efficient basis.

The first recorded multiple method of preparing splints was that patented by Phillips in 1836 (*1*). Phillips cut thin slabs of pine which he further reduced to match length. The slabs were then crosscut by saw gangs lengthwise with the grain of the wood, comb fashion, leaving a portion at one end uncut, holding the splints (generally twelve in number) together like the back of a comb. These were dipped in molten sulfur, then in composition, and packaged for sale by laying the slabs upon long strips of paper cut sufficiently wide to lap over the ends of the matches; the slabs were then doubled up in the paper like a paper of pins.

Another effort toward multiple dipping of splints was that patented in 1837 by Stephen Dole, who pasted a piece of cloth previously sized upon the bottom of a block of dry pine wood (*2*). He then grooved the top end of the block with saws to a depth of about one-half inch, in two directions, the resulting splint cross section being about one-eighth inch square. The block was then split with the aid of a knife or chisel at the sawed incisions, the resulting splints remaining attached to the cloth. When dipped, the free ends of the adhering splints were sufficiently separated to receive the composition, with a minimum number of double-headed matches resulting. Matches of this type may be considered the forerunner of the block match.

In the same year, a method of dipping loco-foco matches was patented by John Hatfield (*3*). According to him, the general practice in dipping matches at this time was to take 20 or 30 splints in the hand, adjusting the bottoms so that they ranged in a somewhat convex line; then by turning the hand, to separate the individual splints at their ends and dip them in composition. Hatfield proposed the use of a conical-shaped metal tube whose sides diverged at an angle of about 30°. A bundle of splints was inserted in the mouth of the tube and pressed tightly until they met the sides of the tube. The converged splint ends were then dipped, being "regularly and efficiently separated" so that few double heads were formed.

Machinery soon replaced many of the hand operations previously required for splint making, however, and in 1840 Winans and Hyatt patented one of the first successful splint-cutting machines (*4*). This machine consisted of two rollers provided with a series of circumferential grooves and mounted so that the grooves coincided, the resulting opening being of the desired size or cross section of the splint. Veneer sheets of wood, when fed into the grooves between the manually operated rollers, were compressed and cut into splint form by the sharp edges of the grooves.

In 1842 C. E. Warner patented (*5*) the first automatic splint cutter designed to use the hollow cutting die—the forerunner of the subsequent automatic cutting head developed many years later and universally used today for round splint matches. Warner's machine consisted of a series of "cutting cylinders" or dies, which automatically operated on a block of wood in the direction of the grain, with the cut splint passing up through the die. On the next cutting stroke, the previously cut splint was forced out of the die, falling into a tray. Warner's cutting head included two series of eight dies each.

In 1845 Fessenden and Knight (*6*) improved on Warner's invention by developing an automatic machine consisting of a series of cutting dies which cut the splints from wood blocks and then transferred them to perforated wooden dipping frames. These were made of wooden slabs provided with vertical and parallel ranges of holes bored through the slabs transversely and at equal distances apart. This principle is used today.

In 1848 L. Smith invented a cutting machine (*7*) with the first recorded automatic method of advancing the wood blocks periodically with each movement of the cutting knives, as is done today.

Continuous Match Machines

The first continuous automatic match machine was patented in 1851 by I. H. Smith of Wolcott, Massachusetts (*8*). Blocks of wood were first cut into plates or veneers of match splint thickness and of double splint length, the grain of the wood running with the measured double length. These plates were then introduced into the match machine, where they were driven forward by feeding rolls operated by a ratchet geared to the driving shaft. Between motions of the ratchet feed, a vertical guillotine knife cut the splints, which were then carried between two belts, the splint ends projecting beyond the sides of the belts. Composition was applied to the splint ends by means of two vertical rotating wheels which dipped into composition reservoirs. The dipped matches were then carried by the belts a sufficient distance to allow the composition to acquire an initial "set," after which the double splints were cut in two and the matches boxed.

A more practical and successful continuous machine was made in 1854 by Gates and Harwood, which may be considered as the first of the successful continuous

[1] Continued from the July, 1941, issue.

and automatic machines developed principally between 1870 and 1880 (*9*). Blocks of wood were fed into the machine; the splints were cut on the downstroke by a series of cutting dies operating with a vertical reciprocating motion, then were carried by the dies on the upstroke into the open jaws of a series of clamps. These were closed by the operation of a cam, the splints being firmly gripped at their ends. The clamps were attached to an endless chain which moved along with an intermittent motion in balance with the operation of the cutting dies. The individual splints were carried along by this endless chain over a sulfur dip tank, where the chain was depressed by rollers so that each splint received the requisite amount of sulfur. The process was repeated for the application of composition. The dipped matches were then forced out of the clamps with the aid of a hammer plunger and fell into a container, ready for packaging.

Increased use of wax matches or vestas also resulted in the development of machinery in order to increase production and decrease cost. A continuous machine for making vestas was patented by Elkan Adler in 1854 (*10*). Adler's machine consisted of a large cylindrical drum around which were wound 14 continuous lengths of wax wicking. The wicking was fed through grooved rollers and channels, the ends entering separate perforations in a dipping board provided with 11 rows of 14 holes each. A knife automatically cut the wicks to splint length and another perforated board moved up to take the place of the one previously filled. The boards, each of which contained 154 matches, were then dipped by hand. The inventor claimed that one person could prepare and dip 50,000 boxes of matches per day with the use of his machine.

Other so-called automatic or continuous machines patented prior to 1860 include those of Thomas (1855) (*11*), Underwood (1856) (*12*), Holms (1856) (*13*), and Cook (1857) (*14*). There is little evidence to show that these early machines were successfully adapted to large scale production. The large number of relatively small establishments (75 in 1860) and the availability of cheap labor, together with the inherent limitations of the inventions themselves, restricted their use to minor applications only. This is confirmed by the following statement taken from the patent specification on coil dipping (*15*) which was granted to Anson and Ebenezer Beecher on February 3, 1863:

"Prior to our invention, matches have been dipped in two ways, *viz.*, the bundle dip and the frame or clamp dip. In the former case the dipper takes a bundle of splints tied around with a string and grasping it with both hands, causes the ends of the splints, by a skillful movement, to expand into so wide a circle that they become detached, and upon being dipped, each will receive its separate portion of the composition. In frame or clamp dipping as practiced before our invention, the splints are placed by hand or by machinery between members of a series of clamps arranged in a frame, the grasping surfaces of the clamps being grooved at suitable intervals to separate the splints properly, or the splints are set by hand or by machinery in a series of parallel holes through a board of suitable size for a dipping frame. Endless bands have been used to connect and carry a series of clamps fastened thereto instead of being secured to a frame as above mentioned."

That the bundle dip method was in use even as late as 1867 is shown by a patent taken out in that year by William Baustian (*16*). He proceeded in the following manner:

"I take a bundle of sticks of the usual size containing 500 or 1000, tie them together with a string near the tail end, then put the head end of the bundle upon a table or board or upon the palm of my left hand, and press the palm of my right hand upon the tail end of said bundle; then I turn the palm of my right hand a little around, whereby the heads of the sticks are turned so far apart or separated that they can be properly dipped; then after the bundle has been heated upon a stove, I dip it into liquid sulphur. This sulphur dries forthwith on the sticks, and then I rub the bundle in order to remove the clodded sulphur between the matches and then put them back to their original position in the bundle. Then I turn the matches again by the palms of the hands or otherwise, as above described, so that the heads be set apart, and dip them in composition. Instead of cutting the sticks at the usual length, cut them at double length and proceed for one end as for the other, after which, both ends being dipped, cut the sticks in the middle."

That Baustian was a match maker of somewhat reactionary tendencies is indicated by his suggestion that by accurately applying his process, all machinery then in use (1867) for dipping matches could be dispensed with. It also gives a fair indication of the ineffectiveness of the match machinery then in use.

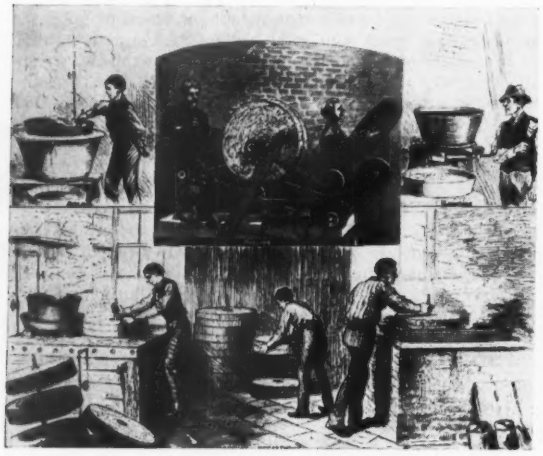
In 1863 Bell and Higgins introduced (*17*) an improved dipping frame mechanism, whereby the frames moved along a track with the aid of rollers, the matches receiving the composition dip by passing over a roller provided with a series of circumferential grooves—one groove for each longitudinal row of splints. This is the first recorded use of the grooved roller for dipping splints mounted in perforated plates or frames.

Coil Dipping

On February 3, 1863, a patent was granted to Anson Beecher and his son Ebenezer, for an "improved machine for framing lucifer-match splints for dipping" (*15*). This process became popularly known as the "coil" method, and was immediately successful. Before the method became out-moded and obsolete because of the invention of the modern continuous machine by the same Ebenezer Beecher in 1888, it became widely adopted by the larger match manufacturers, and it may truly be said that it was the first practical, large scale, mass production method of match manufacturing in the history of the industry.

The original Beecher machine consisted of a cylinder, grooved across its periphery with a regularly spaced series of grooves; a splint-feeding mechanism and hopper designed to provide the cylinder grooves with individual wooden splints; and a long flexible tape of cotton cloth with suitable winding mechanisms. The rotating cylinder was designed to place the splints crosswise upon the tape at regular intervals, the tape then being wound up on a metal core into a large circular coil approximately two feet in diameter which held several thousand splints. The splints were of double length and were thus dipped at both ends. The dipper then removed the coil from the machine,

inserted a handle through the core, and dipped the splints projecting from each face of the coil first in sulfur, and then in composition. The coil was then hung on a rack until the matches were dry enough to permit the unwinding of the coil.



From original drawing by C. A. Keelles in the June 22, 1878, issue of "Harper's Weekly"

BEECHER'S COIL DIPPING PROCESS

Showing method of mixing and grinding the phosphorus composition; rolling the coils; dipping coils in molten sulfur (left) and in composition (right)

The coil method was described in an article written in 1878 as follows (*18*):

"The modern process is an elaborate one. The wood (white pine) comes mostly from Canada, our southern pine being too resinous and cross grained for use. The wood is sent here cut in pieces of proper thickness and twice the length of a match. The first process, 'rolling,' consists in building up the splints of wood by means of belting, so that a compact, circular mass is formed having a diameter of about two feet and a thickness of five inches, the final roll appearing not unlike a circular grindstone. This is done by the rolling machines. The belting is in two rolls, which are placed on axles. From one of the rolls, belting is passed over a series of wheels and around a cylinder under the feeder; its end is then fastened to an axle near the feeder. This feeder contains the splints.

"From the other roll, belting comes to the same axle more directly. When the machine is in motion the belt passes under the feeder; the splints fall on it at right angles and are carried along to the axle where they are caught between the two belts. The arm to which the axle is attached rises as the size of the roll increases, so that the splints always come into contact with the roll at the same distance from the feeder. When the roll is sufficiently large, one or two turns of the belting are made around it and the ends fastened by driving a nail through them into the bundle. The roll is then taken to be dipped into a steam-heated pan of melted sulphur. The roll is dipped, first on one face and then on the other, allowing a short time between dips for partial drying."

According to the author, phosphorus composition was then applied in similar manner, the composition being spread out over a wide-surfaced table with a water-jacketed iron top which was kept heated to maintain the composition in a workable consistency. The rolls were then hung on pegs to dry, and after drying were unrolled on a machine which deposited the splints in orderly fashion on trays. The double-length matches were then cut in two with the aid of a pedal-operated guillotine knife, after which the matches were assembled and boxed by hand.

The coil method was further improved by Scanlon's patented machine for dividing the finished double matches, in 1873 (*19*), and by Bonhack's patent for unwinding the double matches from the dipped coils, in 1882 (*20*).

Young's Automatic Machine

Although the principles of the grooved dipping roll, the perforated match plate, and the reciprocating cutting head provided with a number of dies acting upon wood blocks had been previously patented, McClintock Young was the first to assemble or correlate these basic principles together and, with original refinements of his own, to unite them into one integral mechanism that was commercially practicable (*21*).

Young's patent was granted in 1871. His perforated plates were made of cheap wood, with five holes to a row, corresponding to the five cutting dies with which the machine was equipped. The cutting dies were arranged in a receding line, and cut the splints from the blocks of wood on the downstroke, the splints passing up through the dies. On the upstroke of the reciprocating head the splints were inserted into the plate perforations, with the aid of suitable carrying guides. The plates were then carried along a track over grooved rollers, which applied first the sulfur, then the composition. The dipped matches were then either pressed from the plates and boxed for sale, or sold as block matches. In the latter case the plates were cut into convenient lengths and packed and sold in this condition.

In 1875, Young's second patent was issued, for a punching-out mechanism which automatically punched the splints from the match plates, the operation being synchronized with the driving mechanism of the match machine (*22*).

In 1877 Young developed an improved type of cutting head, provided with two sets of cutting dies which acted upon two blocks of wood instead of the previous single block (*23*). The block feeding troughs were equipped with ratchet feeders, synchronized with the movement of the cutting head, to advance the blocks at the proper interval of time. In addition, a new method of synchronizing the action of the cutting dies with the movement of both blocks and plates was worked out. Young's improved machine was also the first to be equipped with a suction "wind trunk" to remove splinters from the cutting head in order to avoid interference with the action of the cutting dies upon the blocks.

Other automatic and continuous machines invented during this period but which appear to have had little commercial application were those of Cook and LaBelle in 1883 (*24*); Norris and Hagan in 1885 (*25*); and Levi Kittenger in 1886 (*26*).

Splint Manufacture

An appreciable percentage of the matches made between 1860 and 1886 were produced in match factories from prepared splints purchased from independent splint cutters. In most cases, such splints were square or rhombic in cross section; few of them were round. One method described in 1870 utilized pine planks four feet long and two inches thick (*27*). The planks were fed into a slabbing machine, where a cutting knife cut the planks into slabs four feet long and of match splint thickness. These slabs were then passed under a cylinder equipped with a series of small knives spaced the width of a match splint apart. The knives cut the slabs into thin splints, four feet long, which were sold to match manufacturers in bundles of 1400. The splints were reduced to four-inch length at the match factories by guillotine knives, often operated by children, who worked ten hours a day placing the splints into frames for dipping.

The dipping frames consisted of two uprights secured to a crosspiece on their bottoms, and having grooved sides which held the matches firmly in place, just far enough apart to keep them from touching each other. The splint ends protruded from either end of the frame, and both ends received a sulfur dip, followed by application of composition. When dry, the double matches were removed from the frame, cut in two by a pedal-operated guillotine knife, and boxed. Girls were employed for the cutting and boxing operations; it is stated that they operated on a piecework basis, and were capable of cutting, assembling, and packaging 400 boxes per hour per girl.

Another method utilized planks $15'' \times 3'' \times 11''$, cross-cut into blocks $4^3/_4'' \times 3'' \times 11''$. The blocks were then steamed for twenty minutes to soften them up for the splint-cutting operation, then cut into splints by a series of knives. The double-length splints, $4^3/_4$ inches long, were then dried, assembled into bundles, and sold to match manufacturers, who utilized them directly in their dipping frames, each holding 3900 double-length splints.

Other methods used for cutting splints were those patented by Wallace in 1875 (*28*) and duBois in 1876 (*29*). In 1883 Young patented a method of filling perforated match plates with loose splints (*30*). Prior to this invention, the bulk of the prepared splints were used by those factories operating by the coil-dipping process, although many of them were still being dipped in small establishments with the aid of hand-manipulated dipping frames.

An article published in 1877 gives the following brief account of manufacturing practice in that year (*31*):

"Formerly, splints were made by hand, and they were hand dipped. Separate factories are now used for making splints, and all the labour is performed by machinery. As the manufacturers do not allow visitors in their buildings it is impossible to describe exactly the machinery or the methods employed. By one process, white pine lumber is reduced to 2-inch plank and cut into blocks. These blocks are fed into a machine which cuts 12 splints on each stroke and on the following stroke delivers them to an endless chain from which they are dipped successively in melted sulphur and in composition, then placed in trays for packing."

The method referred to above is that of McClintock Young.

A new form of splint was invented in 1870 by Robinson (*32*), who stated that only two types of splints were in general use up until the time of his invention: square splints, sawn or cut from blocks; and round splints, cut by dies from blocks. Robinson's splints were diamond-shaped in cross section, thereby providing two thin feather-like edges which facilitated the ignition of the splint from the flaming match head. The inventor stated that the two existing types of match splints caught fire only with great difficulty, especially when they were not sulfur-coated. This indicates that at least some of the matches in use at this period were sold without having had a preliminary flame-transmission treatment of sulfur or paraffin.

The principle of Robinson's invention has been carried over into the present day. Practically all round splint matches are provided with splints which are not truly cylindrical in shape, but which have two or three small concave indentations running lengthwise up and down the splint. The edges of these indentations provide the feather-like surfaces which ignite quickly and easily from the flaming head, transmitting the fire from the latter to the treated splint.

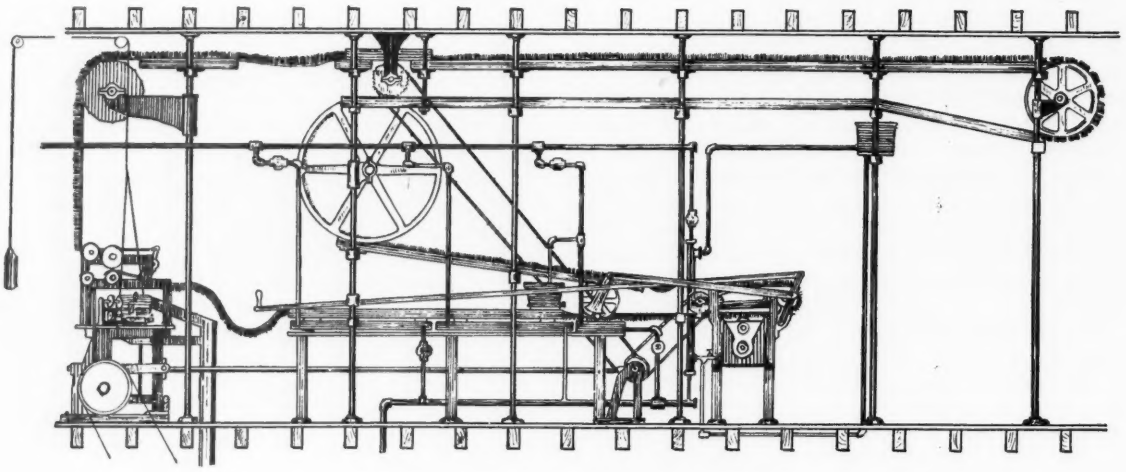
Beecher's Continuous Machine

On September 11, 1888, the co-inventor of the coil-dipping process, Ebenezer Beecher, was granted a patent for a continuous, automatic match machine, which revolutionized the match industry and laid the foundation for the present-day high-speed continuous machine (*33*).

Beecher's match machine consisted of an endless chain arrangement made up of a series of independent perforated metal match plates, linked together at their edges; a reciprocating cutting-head mechanism provided with a series of hollow cutting dies which cut a single row of 36 match splints from a wooden block on the descending movement, carrying these splints up to, and inserting them into the perforations in the match plates on the ascending movement; a heater to heat the splints before dipping them in paraffin; a paraffin dipping tank; a composition dipping tank provided with a grooved roller through which the rows of splints received a uniform coating of composition; and a drying track, over which the filled match plates were conveyed for a sufficient length of time for drying the matches. The plates were then conveyed back to the cutting head, where the individual matches were punched out of the plates and boxed. The cutting head was also provided with a suction system for removing the splinters of wood, in order that the action of the cutting dies on the blocks would not be interfered with.

In 1894 Beecher, in collaboration with J. P. Wright, developed further improvements in the cutting-head

mechanism, cutting dies, driving mechanisms, composition tank, and match-discharging mechanism (*34*). Additional length was given the endless plate arrangement in order to dry the matches more completely before packing them in boxes. These improvements, when added to Beecher's first machine, resulted in what is essentially the modern high-speed machine of today.



E. B. BEECHER'S CONTINUOUS MACHINE (1888)

LITERATURE CITED

(1) PHILLIPS, U. S. Patent 68 (1836).
(2) DOLE, *ibid.*, 214 (1837).
(3) HATFIELD, *ibid.*, 219 (1837).
(4) WINANS AND HYATT, *ibid.*, 1867 (1840).
(5) WARNER, *ibid.*, 2700 (1842).
(6) FESSENDEN AND KNIGHT, *ibid.*, 4013 (1845).
(7) SMITH, *ibid.*, 5496 (1848).
(8) SMITH, *ibid.*, 8066 (1851).
(9) GATES AND HARWOOD, *ibid.*, 10,737 (1854).
(10) ADLER, *ibid.*, 11,405 (1854).
(11) THOMAS, *ibid.*, 12,297 (1855).
(12) UNDERWOOD, *ibid.*, 14,782 (1856).
(13) HOLMS, *ibid.*, 15,238 (1856).
(14) COOK, *ibid.*, 17,263 (1857).
(15) BEECHER, *ibid.*, 37,562 (1863).
(16) BAUSTIAN, *ibid.*, 62,922 (1867).
(17) BELL AND HIGGINS, *ibid.*, 38,878 (1863).
(18) *Harper's Weekly*, **22**, 488–90 (1878).
(19) SCANLON, U. S. Patent 134,565 (1873).
(20) BONHACK, *ibid.*, 254,191 (1882).
(21) YOUNG, *ibid.*, 118,502 (1871).
(22) YOUNG, *ibid.*, 160,376 (1875).
(23) YOUNG, *ibid.*, 187,576; 194,319 (1877).
(24) COOK AND LABELLE, *ibid.*, 277,462 (1883).
(25) NORRIS AND HAGAN, *ibid.*, 332,737 (1885).
(26) KITTENGER, *ibid.*, 341,809 (1886).
(27) WATSON, *Lippincott's Magazine*, 162–7 (Feb., 1870).
(28) WALLACE, U. S. Patent 169,867 (1875).
(29) DUBOIS, *ibid.*, 182,750 (1876).
(30) YOUNG, *ibid.*, 271,563 (1883).
(31) *Popular Science Monthly*, **11**, 413 (1877).
(32) ROBINSON, U. S. Patent 105,727 (1870).
(33) BEECHER, *ibid.*, 389,435 (1888).
(34) BEECHER AND WRIGHT, *ibid.*, 528,457 (1894).

A NEW PRODUCT

REPRESENTING what is described as a new technology in control of filtration and flow of liquids and gases, a porous metal product known as "Porex" has just been introduced by Moraine Products Division of General Motors Corporation. It is manufactured from powdered metal subjected to a series of processing operations.

Applications for "Porex," the announcement said, are found in almost any appliance or piece of industrial equipment involving the flow of gases or liquids. These uses are found in pumps, refrigerators, fuel lines, lubricating systems, oil burners, evaporators, absorbers, paint sprayers, premixed gas burners, and other devices.

This new filter and diffusing material, for example, is used to prevent clogging of orifices in Diesel injector nozzles with effective removal of fibrous materials not stopped by other types of filters. This new type of filter also acts effectively in preventing a drying agent from passing from its chamber into a refrigeration system when the refrigerant passes through it to be dried.

One important and unique feature of "Porex" is its ability to be bonded to steel and copper. In this way, "Porex" can be made an integral part of solid metal which may be machined, ground, bored, threaded, or processed for individual requirements.

When used as a filter, "Porex" provides conditions which permit high flow rates, low flow restriction, and a high degree of surface clogging before efficiency is reduced. Utilized as a diffusing plate, "Porex" offers low flow restriction and uniform, controlled porosity.

Advantages of Semimicro Technic in Teaching Qualitative Analysis[1]

PAUL ARTHUR, J. A. BURROWS, O. M. SMITH, and EVERETT L. ADAMS

Oklahoma Agricultural and Mechanical College, Stillwater, Oklahoma

AMONG the most frequent questions asked by prospective users of semimicro technic is the all-important one concerning the probable expense involved in changing from the macro to the semimicro equipment. An unusual arrangement of laboratories in the chemistry department at Oklahoma Agricultural and Mechanical College has enabled this department to keep accurate records, over a period of four years, for the semimicro qualitative analysis course. The analytical procedure used involved the customary hydrogen sulfide separation; consequently, it was possible to get laymen's opinions on the nuisance created by the use of this reagent. The experiences at this college are described here in the hope that the information thus furnished will be useful to others.

The costs for such a course are very conveniently divided into two groups: (1) those which would be necessary in changing from the macro to the semimicro procedure—the *starting costs*, and (2) those necessary to maintain the course—the *current costs*.

The actual costs involved in equipping the laboratory at this college for the course in semimicro qualitative analysis are given in Table 1.

TABLE 1

STARTING COSTS

(for 20 students)

Apparatus kits at $5.60	$112.00
Reagent blocks and vials (5 sets)	26.20
Dropper bottles for test solutions for the side shelf	10.32
Burners	11.00
Centrifuges, two at $17.50	35.00
Total	$194.52
Cost per student	9.78
Macro apparatus (returned to other courses)	386.00
Saving on maintenance of H_2S and hood system	Undetermined

These figures cover equipment for twenty students except as otherwise indicated, and are based on the assumption that the department has none of the equipment already in stock.

To make the picture even clearer, Table 1 includes also the cost of equipping an equal number of students with the apparatus which was used at this college when the course in qualitative analysis was taught by macro methods. When the change in methods was made, all of this macro equipment was absorbed in other courses; consequently, the change actually resulted in a saving of over $190 for each twenty students, during the first year.

The second cost question, which is of even greater importance, is that of the running expenses. Careful accounts, covering eight semesters, with an average of 40.6 students taking the course each semester, give the average results tabulated in Table 2.

TABLE 2

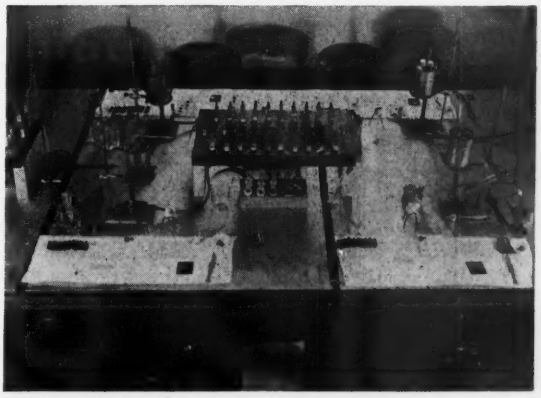
CURRENT COSTS PER SEMESTER

(for 20 students)

Reagents and stock solutions	$5.40
Laboratory upkeep charge	1.40
Student breakage and non-returnables at $2.23	44.60
Total	$51.40
Cost per student	2.57

These figures, again, are based on twenty students. It will be noticed that the student breakage and the non-returnable item is the only costly one in the list. The current laboratory costs, exclusive of gas, water, and electricity, average only $2.57 per student for the whole course.

Though the question of expense is quite important, no discussion of semimicro methods of teaching qualitative analysis is complete without considering some of the other features of such a course. A description of the arrangements used at this college again will serve to bring out some of these features. The qualitative analytical laboratory here is located in a large room on the third floor of a building, the remainder of which is occupied by the *music department!* The equipment in



EQUIPMENT FOR SEMIMICRO QUALITATIVE ANALYSIS

the room is of surprising simplicity. There is a long trough along the side of the room with water faucets along the trough to permit its use for washing apparatus. Plain tables with tops 4 feet by 20 feet, and of just sufficient height (30 inches) to permit the students

[1] Much of the material in this article was presented in a paper before the Division of Chemical Education at the 101st meeting of the A. C. S., St. Louis, Missouri, April 9, 1941.

to *sit on ordinary chairs* and work comfortably, serve as laboratory desks. The tables are equipped with gas, but not with water. The view across the table is unobstructed, since only large, flat, reagent blocks, each containing seventy reagents in vials, are on the tables. The liquid reagents and solutions are in 3-dram vials equipped with Barnes' type droppers; the solids are in one-dram vials fitted with ordinary stoppers. Each of these reagent blocks serves four students, two on each side of the table. Acids and bases are in one-ounce bottles in smaller blocks, two of these being placed by each large block. Each student is provided with a hydrogen sulfide generator of the Aitch-Tu-Ess type. Test solutions and a few of the less stable reagents are the only chemicals not within easy reach of the *seated* student; consequently, moving around the room is reduced to a minimum. There are neither hoods nor special ventilating equipment in the room; yet the other classes in the building have, as yet, made no objection to the practice of the usually very objectionable qualitative analysis. Two electrically driven centrifuges[2] and a large bottle of distilled water complete the equipment of the laboratory. The maximum capacity of this laboratory is thirty students, in a space which would ordinarily have accommodated not more than twenty students using the macro technic and apparatus.

The teaching value of semimicro methods has been pointed out by many teachers of qualitative analysis, but the students themselves are the best source of information on this subject. Conversations with students, two or three years after they had taken this course, have indicated that the great respect for cleanliness and careful technic taught in the semimicro course made them feel that later courses were sloppy and clumsy in comparison. The realization that bulk is not important is one thing that seems surprising and interesting to most students. The course is quite self-motivating in practice; and requests from students who have had the macro course for permission to take the laboratory part of the semimicro course are by no means uncommon.

[2] These centrifuges may be obtained from W. Edmund Fish, P. O. Box 414, West Lafayette, Indiana.

The Effect of Moisture on Chemical Reactions[1]

GEORGE CHAMLIN

Herzl Junior College, Chicago, Illinois

THAT the presence of moisture is necessary for certain chemical changes to occur has been known for a long time, but it is comparatively recently that a detailed investigation along this line has been made. The discoveries of Baker whereby he showed that the most minute traces of water influenced the physical properties of substances have pushed this field to the forefront of chemical research. Although a satisfactory explanation can be offered for the presence of water as a requisite for the occurrence of certain chemical changes, there are still a far larger number of reactions for which no explanation is available at this time. A vast unexplored field lies before the research worker who is interested in this subject.

It is best to state at the beginning that chemical purity is a relative term and absolute dryness is impossible in the laboratory. Dittmar and Henderson found that air dried by calcium chloride retains 0.001 g. of water vapor per liter (*1*). Dibbits found that 308 l. of air dried by means of sulfuric acid gave up 0.07 mg. of moisture using phosphorus pentoxide as a desiccant (*2*). Morley found that 40,000 l. of gas dried by phosphorus pentoxide still retain one mg. of moisture (*3*).

It is well known that lead acetate paper has to be moistened before hydrogen sulfide gas will react with it; that dry hydrogen chloride does not redden blue litmus. We assume that the reason no reaction takes place in the dry state is that dissociation of the reacting substances into active ions cannot take place in the absence of an ionizing medium (water in this case).

In the year 1794 came the very interesting work of Mrs. Fulhame (*4*). She found that if moist pieces of silk impregnated with gold salts were exposed to the action of hydrogen the salt was reduced and the material became gilded. If the reacting substances were exposed in a dry state the reduction did not occur. This was found to be a general property of cloths impregnated with salts of silver, gold, mercury, copper, lead, and antimony, and of reduction by means of hydrogen, phosphorus, sulfur, alkaline sulfides, hydrogen sulfide, phosphine, and methyl iodide. She observed that ethyl alcohol and ethyl ether had not the same effect in the promotion of reduction. She concluded that the presence of moisture did not have the effect of breaking up the salt into particles nor of condensing the gas and so bringing the hydrogen into closer contact with the metallic oxide. She supposed that there must be a double reaction, the hydrogen gas uniting with the oxygen of the water, while the hydrogen of the water in its nascent condition unites with the oxygen of the oxide, reducing it and forming water according to the following equations:

$$H_2 + HOH \rightarrow H_2O + 2(H)$$

$$2(H) + AuO \rightarrow H_2O + Au$$

Parnell showed that hydrogen sulfide in the dry state did not react with papers impregnated with salts of

[1] Presented before the Student Associate Group of the Chicago Section of the A. C. S., Chicago, Illinois, January 24, 1941.

lead, mercury, and copper (*5*). Following are his observations:

(1) Reacted only on particular salts.
(2) To restore reaction water may be present in a state of combination with the salt (hydrate) and can then exert no solvent power.
(3) Moistening salts with absolute alcohol, which dissolves six times its volume of hydrogen disulfide, and exposing to hydrogen disulfide, still gave no reaction.

The reaction between chlorine and the salts potassium iodide and potassium bromide will serve to bring out the quantitative relationship between moisture and rates of chemical reactions. Parsons (1925) found that dry chlorine gas had no appreciable effect on the weight of potassium iodide at room temperature (20°–24°C.) even after passing chlorine gas over it for more than one hour (*6*). The only evidence of any chemical reaction is that the iodide turns a pale straw-yellow. On performing a similar experiment with KBr in carbon disulfide medium, I observed a yellowish color imparted to the carbon disulfide. On the addition of a drop of water there is a liberation of heat at the point of contact of the water and the potassium bromide crystals, and the carbon disulfide assumes the orange-reddish color characteristic of molecular bromine. Table 1, worked out by Parsons, shows the effects of various concentrations of water vapor in chlorine on the extent of conversion of potassium iodide crystals (*7*). In each case the chlorine was passed for an hour at the rate of one hundred bubbles per minute at a temperature of 20°–24°C.

TABLE 1

Partial Pressure of Water Vapor (mm.)	*Conversion of Iodide, %*
0.0	0.04
4.2	0.4
7.3	1.7
10.0	3.6
14.2	5.6
23.5	10.7

The purified iodide was dried at 120°C. and then powdered while still hot. It was kept in a desiccator over phosphorus pentoxide until used. Chlorine gas was dried by two towers of sulfuric acid and then by four tubes of phosphorus pentoxide. The extent to which the reaction proceeded was measured by obtaining the loss in weight of the well-dried halide, after the iodine set free by the chlorine had been removed by evacuation and by heating to 200° to 250°C. in a current of dry air. The partial pressure of water in the chlorine for amounts up to 10 mm. was controlled by bubbling gas through different concentrations of sulfuric acid.

TABLE 2

Temperature (°C.)	*Partial Pressure of Water Vapor (mm.)*	*Conversion of Bromine, %*
20	6.0	0.5
	10.1	1.6
	12.0	12.5
	14.2	14.5
	17.4	20.3
30	14.2	0.5
	19.6	1.7
	22.1	9.6
	25.0	11.5

Similar results were obtained using potassium bromide crystals (Table 2).

The sudden increase at 12.0 and 22.1 mm., respectively, coincides with the partial water vapor pressure of the saturated salt solution. Hence the results seem to indicate that the reaction takes place only in solution. The effect of water in this case is perhaps due to a double reaction.

$$Cl_2 + HOH \rightarrow 2HCl + (O)$$
$$H_2O + O + 2KI \rightarrow 2KOH + I_2$$

I believe that it is not the activity of the halogen in the electrochemical series, but the nascent oxygen which causes the displacement. Since this reaction does not take place in the absence of water it may be due to the lack of solution and breakdown of the crystal lattice of the halide. If a substance could be had with a dielectric constant approaching that of water and which would not liberate nascent oxygen it would be possible to find out if lack of solution of the salt is the cause.

A few of the arguments set forth against the belief that water is necessary in order for chemical reactions to occur are those given by Hinshelwood (*8*):

1. No homogeneous gas reaction with a simple unimolecular mechanism is stopped in the absence of water vapor. An example of this is the reaction between hydrogen and iodine.

$$HI + HI \rightarrow H_2 + I_2$$

This reaction varies with the temperature and the pressure. Presence or absence of water does not affect the result.

2. Direct disproof cannot be had; the claim can be made that the substances were not pure or dry enough.

3. Indirect disproof; he claims that the observed rate of reaction is very nearly the same as the maximum possible rate. If the reaction could take place only when two molecules met in the presence of a third substance (catalyst), present in minute amounts, the observed rate would be smaller than the maximum possible. If chemical combination of elements could not take place at all in the absence of water, all elements in the universe would be free today, since hydrogen and oxygen would never have combined and made other combinations possible.

The theories set forth to explain the effect of water upon chemical reactions are numerous and varied.

The Intermediate Compound Theory was proposed by William Higgins in 1814 (*9*). Let us consider the reaction between hydrogen sulfide and iodine in ether solution. The reaction proceeds according to this equation:

$$H_2S + I_2 \rightarrow 2HI + S$$

If the reaction takes place in an aqueous solution it proceeds by these steps, according to Parsons (*10*):

$$3H_2O + 3I_2 \rightarrow 6HI + 3(O)$$
$$3(O) + H_2S \rightarrow H_2SO_3$$
$$H_2SO_3 + 6HI \rightarrow 3H_2O + H_2S + 3I_2$$

The overall reaction is:

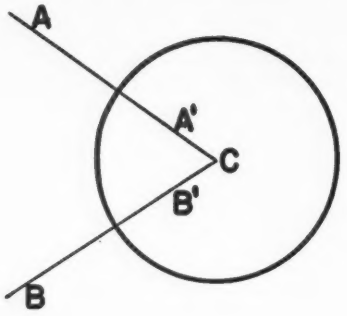
$$H_2S + 3H_2O + 3I_2 = H_2SO_3 + 6HI$$

The side reactions have already been mentioned. The greatest objection to this theory is that if water

does form an intermediate compound or enters into a side reaction, the concentration of the water vapor should be important. Coehn and Jung found that when the vapor pressure exceeds the very small value of 10^{-3} mm., its concentration has no effect whatever on the rate of the reaction between hydrogen and chlorine.

Solution Theories. Magnesium and iodine will not combine in a dry condition. If, however, a drop of water is added a violent reaction occurs and synthesis takes place. This reaction also takes place if ether, pyridine, alcohol, or acetone is substituted for water. In those liquids which do not bring about reaction (carbon disulfide and carbon tetrachloride), magnesium iodide is not soluble. The explanation is that a thin film of magnesium iodide is formed on the reacting substances, preventing further chemical reaction. This film is too thin to be measured and can be compared to the oxide film on aluminum metal. Water and other liquids apparently wash away this film by mechanical action or by forming complexes with the liquid and thus exposing fresh reacting surface (*11*).

The Electrical Force Theory was developed as a result of the experiments which Sir J. J. Thomson had carried out showing the great influence of water vapor on the passage of electrical discharges through gases (*12*).



Suppose the forces holding the atoms in the molecule to be electrical in character. This force will be very much diminished when the molecule is close to a conductor or a substance like water possessing a very large specific inductive capacity (dielectric constant). Let A and B represent two atoms in a molecule placed near a conducting sphere. The electrical charge induced on the sphere by A will be represented by an opposite charge at A'. If A is very near the inductive sphere then the opposite charge will almost be equal to A, neutralizing A. The affinity between the two atoms A and B will be almost annulled by the presence of the sphere. Thus, the atoms will be free to take part in chemical reactions.

Smits' Theory (*13*) supposes that every substance contains at least two different kinds of molecules which are in equilibrium. Only one of these forms is chemically active.

$$NH_{3\alpha} \leftrightarrows NH_{3\beta}$$

He suggests that on intensive drying this inner equilibrium is shifted toward the inactive side. This theory seems to have a wider application to chemical reactions than any known.

A classification should now be made of reactions in which water seems to play an important role; those requiring rough drying and those of higher desiccation. We know that a drop of water is necessary for a reaction to occur in baking powder, the water causing hydrolysis. The reactions of a more or less ionic character fall into this group. The absorption of carbon dioxide by alkaline hydroxide requires the presence of an ionizing medium before the ionic reactions can take place. Hence even rough drying will prevent the reaction from progressing to any measurable extent. At the other extreme we have reactions such as those between hydrogen and chlorine and between hydrogen and oxygen. In these cases higher desiccation is necessary to prevent reaction.

LITERATURE CITED

(1) DITTMAR AND HENDERSON, *Proc. Roy. Soc.* Glasgow, **22**, 33 (1891).
(2) DIBBITS, *Z. anal. Chem.*, **4**, 180 (1876).
(3) MORLEY, *Am. J. Sci.*, (iii) **30**, 140 (1885).
(4) FULHAME, "Essay on combustion, with a view to a new art of dyeing and painting wherein the phlogistic and antiphlogistic hypotheses are proved erroneous," **1794.**
(5) PARNELL, *Brit. Assoc. Advancement Sci., Rep.*, **1841**, p. 51.
(6) PARSONS, *J. Am. Chem. Soc.*, **47**, 1817 (1925).
(7) *Ibid.*
(8) HINSHELWOOD, *School Sci. Rev.*, No. **31**, 169 (1927).
(9) HIGGINS, "Comparative view of the phlogiston theory," **1791**; "Experiments and observations on the atomic theory," **1814.**
(10) PARSONS, *J. Am. Chem. Soc.*, **47**, 1820 (1925).
(11) Cited by SMITH, "The effects of moisture on chemical and physical changes," Longmans, Green and Co., Inc., New York City, **1929**, p. 121.
(12) THOMSON, *Phil. Mag.*, (V) **36**, 320 (1893); *Brit. Assoc. Advancement Sci., Rep.*, **1894**, p. 482.
(13) SMITS, *Proc. Koninkl. Akad. Wetenschappen Amsterdam*, **26**, 266 (1923).

Much of the world's sorely needed shipping is now devoted to gathering from the four corners of the earth the strategic metals which will influence the outcome of the current wars. To those responsible for raw-material procurement this scattering of civilization's necessities may seem most regrettable. However, the pitifully sparse deposits of metals which man must first discover, usually in the wilderness, and then mine, smelt, and refine at great length, in reality represent an enormous degree of natural concentration from the minute percentages existing throughout the earth's crust. This concentration is due to geochemical processes which act in rocks, water solutions, and living organisms to produce our varied world from the originally homogeneous globe.

—From the *Industrial Bulletin* of Arthur D. Little, Inc., June, 1941

Seventy-Five Years of Chemistry at the University of New Hampshire

JAMES A. FUNKHOUSER *University of New Hampshire, Durham, New Hampshire*

THE Morrill Act of 1862 granted land in the public domain to the several states to found and support colleges for instruction in agriculture and the mechanic arts. The New Hampshire State Legislature passed an enabling act in June, 1866, which accepted the provisions of the Morrill Act and led to the establishment of the original college in loose association with Dartmouth College at Hanover. The University of New Hampshire, arising from these enactments, is therefore celebrating the seventy-fifth year of its existence. Inasmuch as the first faculty member appointed to the New Hampshire College of Agriculture and the Mechanic Arts was the Professor of General and Applied Chemistry, the present department of chemistry has been interested in tracing its history through the intervening years. The study falls naturally into periods dominated in each case by an outstanding personality.

Ezekiel Webster Dimond,[1] a graduate of Middlebury College in 1865, the first faculty member of the institution, *was* the institution in large measure during its early life. Not only was he Professor of General and Applied Chemistry from 1868 to his untimely death in 1876 at the age of forty, but he was also business manager, architect, supervisor of construction, farm manager, and lobbyist in the legislature. Recognizing the need for an experimental farm, he purchased 135 acres with his own funds, since none were available from the college for this purpose. In his first report to the legislature, Dimond stressed the need for an instructional building and in a second report a year later the need for a dormitory. Funds were shortly made available through the combined contributions of Dartmouth College, the State Legislature, and two generous friends of the school, David Culver and John Conant. In the instructional building, Culver Hall, space was provided for a chemistry laboratory, which in 1874 was described as the only one in the state.

While this material growth was being attained, the enrolment had increased from eleven in 1870–71 (3 senior, 3 middle, 5 junior) to twenty-nine (7 seniors) in 1876–77. Corresponding additions to the teaching staff were made, although in this period most of the faculty divided their time between the agricultural college and Dartmouth. The customary stipend was about two dollars per hour of teaching. The range of subjects which could be taken and the importance of chemistry at this early time are shown by the list of studies, as given in the report of the examining committees for the year 1875–76.

[1] WALKER, "Memorial sketch of the life and character of Ezekiel Webster Dimond," Edward Jenks, State Printer, Concord, N. H., 1877.

EZEKIEL W. DIMOND
PROFESSOR OF CHEMISTRY 1868–76

Studies	*Number of Recitations*
Bookkeeping	45
Algebra	164
Botany	76
Physiology	36
Geometry	154
Zoölogy	45
Natural Philosophy	76
Entomology	95
Chemistry	195
Astronomy	70
Meteorology	45
Trigonometry	36
Surveying	30
Geology	54
Political Economy	36

This report marked the end of the first period which was dominated so largely by Dimond's dynamic personality. In honor of his success in guiding the New Hampshire College of Agriculture and Mechanic Arts during its early years, Dartmouth College awarded Dimond an honorary A.M., shortly before his death in the summer of 1876. In the words of President Asa Dodge Smith,[1] "his death was undoubtedly the result of overwork," which was "twice the amount of work of one man."

Looking back on his undoubted success, it is interesting to record that most of the letters of recommendation submitted at the time of Dimond's appointment

were unfavorable. This brings to mind the similar lack of faith in Berzelius' future success as a chemist, which his examiners expressed after passing him somewhat reluctantly.

CULVER HALL
First classroom building of the New Hampshire College of Agriculture and Mechanic Arts. Built in 1870, it honored an early benefactor, David Culver of Lyme, New Hampshire

The next year (1877) Benjamin Thomas Blanpied was appointed to the vacant professorship of chemistry. He was a student of medicine at Dartmouth when first appointed as a part-time instructor in 1871. He never completed his medical studies but many years later (1897) he received the Ph.D. from Bethany College, his *Alma Mater*. While Blanpied did not possess all the remarkable traits of his predecessor, the chemistry department continued to grow. Of the class of thirteen graduated in 1877, one, George M. Holman of Fitchburg, Massachusetts, presented a thesis on "The Chemistry of a Piece of Granite." Certainly there could have been no dearth of samples of this material around Hanover. The class of 1885 produced the college's first professional chemist, Allen Hazen of Hartford, Vermont, who became chemist of the State Board of Health at Lawrence, Massachusetts. In 1886 James C. Harvey of Surry presented a thesis on "Practical Applications of Organic Chemistry." Harvey continued the study of chemistry at the Massachusetts Institute of Technology, thus becoming the first of the department's long line of graduates who have taken advanced degrees. Also in the 1886 class was the first foreign student, Belezar S. Rouevsky of Bulgaria. In 1888, George E. Porter, also of Hartford, Vermont, received the first Bailey Prize, founded by C. H. Bailey '79 and E. A. Bailey '85, for proficiency in chemistry. This is the oldest prize which is given by the University at the present time.

BENJAMIN T. BLANPIED
PROFESSOR OF CHEMISTRY 1877–89

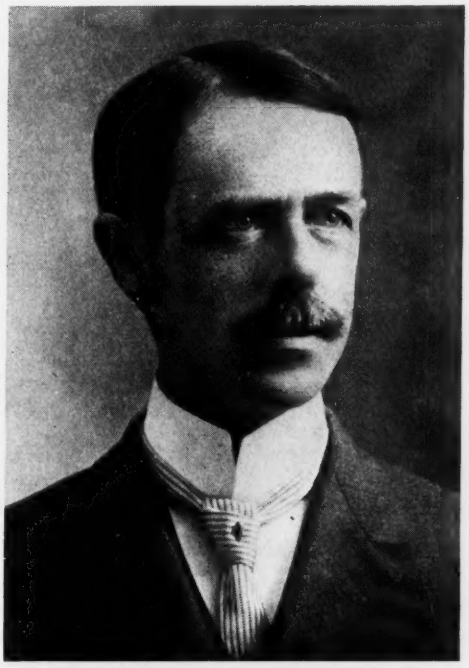
CONANT HALL
This building, erected in Durham in 1893, was the second building to honor John Conant of Jaffrey, New Hampshire, another early benefactor of the college. It originally housed several other science departments in addition to chemistry

The commencement exercises of 1888 included the laying of the cornerstone for the Agricultural Experiment Station building. The staff in this first year all ranked as assistant chemists. They were Edward H. Farrington; Fred W. Morse, Worcester Polytechnic Institute '87; and Charles L. Parsons, Cornell '88. When at the close of the year 1888–89, Professor Blanpied resigned to return to Bethany College, West Virginia, Fred W. Morse became professor of chemistry and Charles L. Parsons instructor. As the experiment station work increased, Professor Morse, who shortly became vice director, found it necessary to relinquish all but about six hours of teaching in analytical and organic chemistry. The other teaching was assumed by Charles L. Parsons, who was made professor of chemistry in 1891. This division of the work continued successfully until Professor Morse left in 1909. This period is marked by the enrolment of the first women students of the college. Miss Lucy Swallow of Hollis

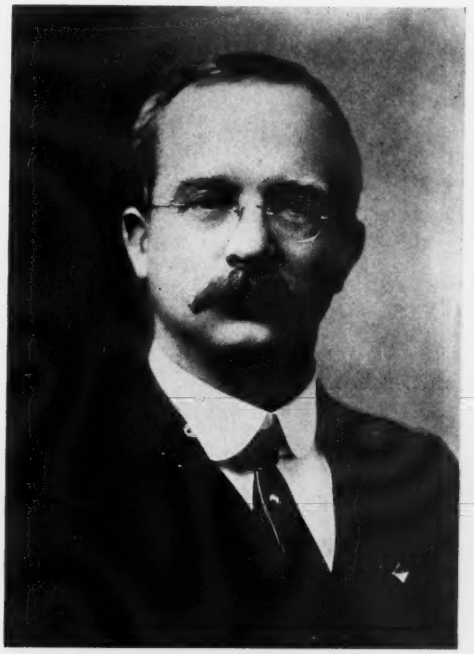
enrolled in 1890 in the chemistry course to prepare for further work at the Massachusetts Institute of Technology. She was joined by Delia E. Brown of Hanover. The college has been coeducational since that time, although few women students have majored in chemistry.

Meanwhile, unknown to the state, there lay in a safe in Durham the will of a farmer, Benjamin Thompson, which was drawn in 1856 and which bequeathed his entire estate to the people of New Hampshire on condition that there would be established on his land a college of agriculture. If New Hampshire failed to accept the provisions of this will, the estate was to be offered to Massachusetts, and if not accepted by Massachusetts the funds were to go to Michigan. Inasmuch as friction between Dartmouth and the College of Agriculture had been developing over various matters, the provisions of the Thompson will, becoming known in 1890, seemed to offer a desirable solution to the problems which had been arising. With the approval of the Legislature the existing College of Agriculture and Mechanic Arts was removed from Hanover to Durham in 1892–93.

By this time the enrolment had reached sixty-four, including ten women. The change of scene gave the college fresh impetus to embark on a period of expansion. Of four buildings erected at Durham, the one occupied by chemistry, and other science departments at first, honored John Conant. This was the home of the department until 1928. While previously occasional pieces of productive work in chemistry had been carried out, Professor Parsons now began to direct the interests of the more advanced students toward systematic research in chemistry. This policy of fostering an interest in research was notably successful and has continued to be an important part of the department's activities in the succeeding years.



FRED W. MORSE
PROFESSOR OF CHEMISTRY 1889–1909



CHARLES L. PARSONS
PROFESSOR OF CHEMISTRY 1891–1911

While its physical growth continued to be slow but steady, the college faculty was building a good reputation for the institution and its products. So in 1909 Professor Morse went to larger duties as research chemist and later acting director of the Massachusetts Experiment Station at Amherst, where he continued until his retirement in 1936. Professor Parsons remained in Durham two years longer, when he too left to become chief mineral chemist of the Bureau of Mines. From 1907 to 1911 Durham was the center of the activities of the American Chemical Society, for in the former year Professor Parsons began his uninterrupted term of office as secretary of that society.

Into the gap caused by these two removals stepped the tall spare figure of Charles James, who became professor of chemistry in 1912. He had come to Durham in 1906 after study with Sir William Ramsay and extensive travel, to begin a remarkable series of researches on the rare earth metals. These researches were in part an outgrowth of Professor Parsons' interest in the rarer metals, beryllium and zirconium, and in part due to Professor James' own interests, fostered by mineral collections gathered during his travels. In

Conant Hall the many thousands of manipulations necessary in making rare earth separations were carried out with painstaking care by Professor James and his students. This work made Durham known throughout

CHARLES JAMES
PROFESSOR OF CHEMISTRY 1912–28

the world as a center of rare earth research. Over the years the escape of hydrofluoric acid during one manipulation etched the window panes so that they came to resemble frosted glass. The last of these have been removed only recently. By his own enthusiasm and interest, Professor James stimulated many students to continue in chemistry as their life work. Many of these men hold positions of importance in the field of chemistry today.

During the war years 1914–18 and the years immediately following, the enrolment grew rapidly so that in 1923 the State Legislature changed the name of the institution to the University of New Hampshire. In chemistry a parallel growth occurred, aided in large part by the attractive personality of "King" James. By 1926 it was apparent that new and larger quarters were needed to house the expanding activities of the chemistry department. During the next two years Professor James spent much time working and planning for the new building. Unfortunately, he was never to work in it, as he died in Boston, December 10, 1928. The new building, named Charles James Hall in his honor, was dedicated on November 9, 1929. The University was joined in the dedication exercises by the Northeastern Section of the American Chemical Society and prominent chemists from other sections. The program was as follows:

Charles James—Teacher.....*Lester A. Pratt* '09
Charles James—Chemist.....*Irving Langmuir*
The Relation of Chemistry to Agriculture and Biology.....
Henry G. Knight
The Laboratory—The Key to Progress.....*Charles L. Parsons*

HAROLD A. IDDLES
PROFESSOR OF CHEMISTRY 1929–

CHARLES JAMES HALL
Erected in 1928, it is the present home of the department of chemistry and chemical engineering and the department of agricultural and biological chemistry

The final period of the department's history began with the selection of Harold A. Iddles as professor of chemistry. A graduate of Michigan State College in 1918, Professor Iddles obtained the Master's degree from the University of Iowa and the doctorate from Columbia. His first task was that of moving from the crowded quarters of Conant Hall to the present building, and expansion of the department offerings. Under Professor James the work in inorganic and analytical chemistry had been of major importance and it was

felt necessary to strengthen the other fundamental branches of chemistry, particularly organic and physical chemistry. This required some expansion of the staff by bringing to Durham some younger men whose interests were in these fields. At the same time this allowed the more advanced students to have a broader choice of fields in which to carry out minor investigational problems. This type of work, which had been started under Professor Parsons and continued under Professor James, was felt to be an important element in the young chemists' training.

Throughout the department's existence its students have received some training in engineering courses, for the department of chemistry is a part of the College of Technology. Consequently the most recent step in development involved the introduction of a chemical engineering option in 1938. This provided two parallel courses, one in chemistry and one in chemical engineering, during the junior and senior years. In addition to the training in the fundamentals of chemistry and chemical engineering, the students are receiving practical training, as they build much of the chemical engineering equipment.

Thus the oldest department of the University of New Hampshire has kept pace with the growth and expansion of the University as a whole. At present the senior staff of seven and junior staff of eight give courses to about 700 students or approximately one-third of the University enrolment. Of these, 120 are majors in chemistry or chemical engineering. While conditions, equipment, and student body are far different from those of Hanover in 1868, no doubt Ezekiel Dimond would regard the present situation with approval, if he were able to return to view the outgrowth of his early labors.

The author wishes to express his thanks to Miss Mildred Saunders, archivist of the Dartmouth College library, for the privilege of examining material in the Dartmouth archives and for the photographs of Dimond and Blanpied.

Derivation of the Freezing-Point Equation

HOWARD M. TEETER

Bradley Polytechnic Institute, Peoria, Illinois

THIS simplified derivation of the freezing-point equation is presented because it is adequately rigorous, it follows the same logic often used in deriving the corresponding boiling-point equation,[1] and in view of its directness, it can be used in briefer courses in physical chemistry to furnish mathematical justification for an equation which is often presented empirically.[2]

The following general assumptions will be made: (1) the solution shall be ideal and dilute, so that Raoult's Law will apply to the solvent and so that the mol fraction of solute, N_2, shall equal n_2/n_1, where n_1 and n_2 are the number of mols of solvent and solute, respectively; (2) the difference, $T_0 - T$, between the freezing points of solvent and solution shall be small so that $T_0T \approx T_0^2$, the vapor pressures of solvent, solution, and solid solvent at the temperatures T_0 and T will be very nearly identical, and the latent heats of fusion, vaporization, and sublimation may be considered constant. These conditions will be fulfilled if $N_2 \leqq 0.01$.

The depression in freezing point, ΔT, is associated with a depression, Δp, in the vapor pressure of the system.[3] This decrease in vapor pressure can be evaluated in two ways. If we follow down the vapor pressure curve of the solid solvent from T_0 to T, we have by the Clapeyron equation

$$\Delta p = \Delta H_s \Delta T / T_0 \Delta V \tag{1}$$

where ΔH_s is the molar heat of sublimation and ΔV is the difference in molar volume of solid and vapor. If we assume the molar volume of solid to be negligible, equation (1) takes the form

$$\Delta p = p \Delta H_s \Delta T / RT_0^2 \tag{2}$$

where p is the vapor pressure.

To evaluate Δp in a second manner we note that it may be obtained in two increments. The first is found by following the vapor pressure curve of supercooled liquid from T_0 to T. The equation obtained is similar to (2) except that now ΔH_v, the molar heat of vaporization, appears:

$$\Delta p_1 = p \Delta H_v \Delta T / RT_0^2 \tag{3}$$

The second increment is due to addition of solute and is given by Raoult's Law:

$$\Delta p_2 = pN_2 \tag{4}$$

where N_2 is the mol fraction of the added solute.

Now since $\Delta p = \Delta p_1 + \Delta p_2$, we have

$$p \Delta H_s \Delta T / RT_0^2 = p \Delta H_v \Delta T / RT_0^2 + pN_2 \tag{5}$$

which reduces easily to

$$N_2 = \Delta H_f \Delta T / RT_0^2 \tag{6}$$

where ΔH_f is the molar heat of fusion. If now we take $N_2 = n_2/n_1$, and $n_2 = w/M$, where w and M are the weight of solute added and its molecular weight, respectively, we obtain (6) in its customary form:

$$M = wRT_0^2 / n_1 \Delta H_f \Delta T \tag{7}$$

[1] *Cf.* MACDOUGAL, "Physical chemistry," The Macmillan Company, New York City, 1936, p. 246.

[2] *Cf.* BIRCHER, "Physical chemistry, a brief course," Prentice-Hall, Inc., New York City, 1940, p. 168.

[3] The depressions are very small, in accordance with the assumptions; however, differentials are not used in order to keep the mathematics as simple as possible.

Apparatus for Producing Liquids of Low Boiling Points

ERNEST B. WILSON *Erasmus Hall High School, Brooklyn, New York*

DRY gases such as methyl ether (b. p. −34.0°C.), chlorine (−34.7°C.), ammonia (−33.4°C.), hydrogen sulfide (−60.3°C.), and sulfur dioxide (−10.0°C.) are easily liquefied by leading them to the bottom of a pyrex test tube surrounded by a mush of solid carbon dioxide mixed with liquids such as acetone, alcohol, ether, carbon tetrachloride, chloroform, or a mixture of equal volumes of chloroform and carbon tetrachloride. Acetone mush foams less; ether mush evaporates more quickly when pressure is reduced; and the mush with carbon tetrachloride, or the mixture with chloroform, is preferable for classroom demonstrations because it is non-inflammable. The container for the mush is a wide-mouthed vacuum bottle of pyrex glass or a beaker wrapped with thick asbestos insulation.

Even hydrogen chloride (b. p. −85.0°C.) and nitric oxide (−89.5°C.) may be liquefied by this method if the mush is under suction from the water pump. The forced evaporation of the liquid, preferably ether, cools the mush considerably below the sublimation point of carbon dioxide (−78.5°C.).

The apparatus mentioned is wasteful of dry ice, is too small for demonstration purposes, and becomes covered with frost which prevents observation of the liquid; also some of the gas escapes into the room to produce annoyance and distraction among the students. To overcome these difficulties the condenser shown in Figure 1 may be employed.

This condenser may be made of pyrex test tubes. The outer jacket should be at least 8 cm. in diameter to prevent frosting. The inner tube should be about 2 or 3 cm. in diameter to permit easy addition of the cooling agent. This tube should extend about two-thirds of the length of the jacket and may be sealed to the jacket if desired. The side tubes attached to the jacket are for rubber tubing, but the bottom tube should be at least one cm. in diameter to permit refluxing of low density liquid ammonia. The length of the condenser may vary depending on the capacity desired. A jacket length of 40 cm. should prove suitable for average use.

An excellent receiver for the condensed liquids is a 30-cc. pyrex test tube held by a stopper in an unsilvered vacuum bottle. The receiver may be placed in a larger test tube or bottle if a vacuum bottle is not available. The outer tube prevents frosting of the receiver when some of the liquid evaporates. A little dry ice may be placed in the bottom of the outer tube if it is of pyrex glass. An opening for the escape of gaseous carbon dioxide must be provided in this case.

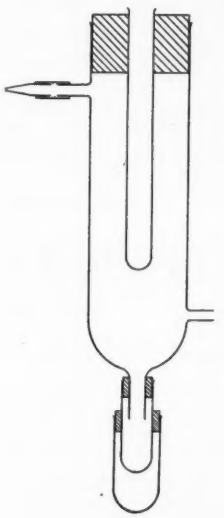


FIGURE 1

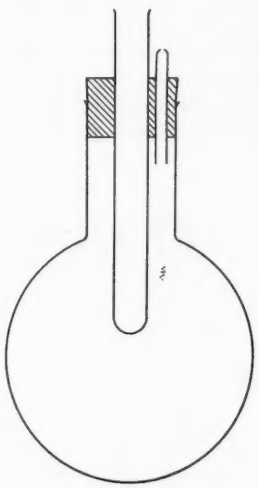


FIGURE 2

The liquids collected may be sealed in tubes for exhibition purposes, used for demonstrating refrigeration, or for special experiments. Crystals may be obtained from the liquids if a few cubic centimeters of each are made to boil by reduced pressure in a test tube surrounded by dry-ice mush. A water suction pump has produced crystals from liquid chlorine (m. p. of pure crystals −101.6°C.), ammonia (−77.6°C.), hydrogen sulfide (−85.5°C.), and sulfur dioxide (−72.7°C.). However, melting-point determinations were not made of the crystals obtained from any of the liquids.

A freezing mixture of calcium chloride and crushed ice instead of dry-ice mush was tried and found satisfactory for gases that liquefy above −20°C. Liquid oxygen caused gaseous chlorine to crystallize on the inner tube in thick layers. Beautiful blue crystals of nitrogen trioxide (m. p. −102.0°C.) mixed with colorless crystals of nitrogen dioxide (−161.0°C.) were obtained also. The crystals melted almost the instant that the oxygen disappeared. When the condenser was filled with gaseous oxygen under slight pressure, a fog of condensed oxygen streamed continuously from the inner tube containing liquid oxygen. No liquid

oxygen was obtained because the fog evaporated as it touched the bottom of the condenser.

This condenser is economical to operate, is large enough for students to see the liquid dropping off the inner tube and collecting in the receiver, and does not become covered with frost. No appreciable amount of gas escapes into the room. Moreover, the apparatus is not expensive or difficult to make. It should become a standard laboratory tool for liquefying gases that condense below 0°C.

However, the production of several cubic centimeters of liquid is undesirable for ordinary lecture demonstrations because of the waste of reagents and the attention required for generating the gas and disposing of the liquid. The reflux chamber shown in Figure 2 obviates these difficulties because very small quantities of reagents are needed, little attention is required, practically no gas escapes, and operation continues as long as desired.

This apparatus has given excellent results when used for sulfur dioxide (b. p. −10.0°C.), chlorine (−34.7°C.), and ammonia (−33.4°C.). After filling with a gas and inserting the dry-ice mush tube, liquid in the form of a fog or large drops begins to fall to the bottom in two or three minutes. In a short time a few cubic centimeters of liquid accumulate. The evaporation of some of the liquid produces a disc of frost on the outside of the flask. A disc of ice may be produced if a watch glass containing about 10 cc. of water is held against the bottom or if the flask is allowed to rest on top of a beaker half filled with water. Wet splints, paper, etc., freeze to the bottom of the flask if held under the liquid. Ammonia produces ice more quickly, chlorine may be seen more easily, while liquid sulfur dioxide is more permanent.

Pyrex chambers varying in capacity from $^1/_4$ liter to 10 liters have been used with success. A large chamber is preferable for lecture demonstrations and small chambers may be used for individual student experiments. The capillary tube extending to the bottom should be bent sufficiently to avoid the pool of liquid when ammonia is used. A 30-cc. pyrex test tube is suitable for holding the mush in the small chamber but may not be sufficient for larger ones. Permanent chambers may be constructed by sealing the mush tube in the flask and filling with gas under slight pressure.

Reflux chambers may be placed in operation in the laboratory for students to study at their convenience. The analogy between the chamber and the modern refrigerator is easy to understand because in both a gas is liquefied and subsequently evaporated to produce lower temperature in part of the apparatus.

We thank Mr. David Soifer and Mr. Julius Zuckerman of Erasmus Hall High School for valuable suggestions and assistance.

LETTERS

The First High-School Chemistry Laboratory?

To the Editor:

The statement of Dr. B. V. B. Dixon (quoted by Professor Harry N. Holmes on page 195 of the JOURNAL OF CHEMICAL EDUCATION for April) that the first chemical laboratory of high-school grade in the United States was probably installed by him in the St. Louis Central High School in September, 1876, would seem to require correction. Laboratory instruction in chemistry was given in the High School of Cambridge, Massachusetts, at least ten years before the date named by Dr. Dixon. Among the chemists who obtained their first instruction in chemistry at the Cambridge High School was the late Dr. Charles E. Munroe, who on several occasions told me of the instruction which he received at this institution between the years 1864 and 1868. I quote from my obituary of Dr. Munroe [*J. Am. Chem. Soc.*, **61**, 1302 (1939)].

"The principal of the school, W. J. Rolfe, the well-known Shakespearean scholar and editor of school texts, was interested in chemistry but instruction in this subject devolved chiefly upon his associate, J. A. Gillet, with whom Rolfe collaborated in the publication of a 'Handbook of Chemistry for School and Home Use' (Boston, 1869). These teachers encouraged the precocious talent of young Munroe by making him an assistant whose duty it was to fill the reagent bottles and look after the chemical supplies."

Additional information about the early chemical courses in the Cambridge High School has been given me by Mr. George W. Rolfe, son of the former principal and a member of the American Chemical Society since 1893. I quote from his recent letter:

"My father William J. Rolfe resigned as principal in 1868 when I was four years old. Mr. Gillet resigned at about the same time and later became the teacher of science (physics and chemistry) in the New York State Normal College (if I recall the name of the institution correctly). I do know that through their influence a large amount of physical and chemical apparatus was bought for the school at that time and also that a chemical laboratory with four or five benches for students was installed, with reagents, probably at the same time, for they were there for the period that I was at school, 1877 to 1881, inclusive, although none of the laboratory equipment was then in use. This I can explain, I believe, for there was at that time a radical change in the policy of the school heads which may have influenced the resignation of my father and Mr. Gillet."

This change of policy, prevalent also in other high schools, is ascribed by Mr. Rolfe to the stress placed by Harvard and other colleges upon classical languages and mathematics for entrance requirements and to the growing number of high-school students who prepared for college:

"The new head of Cambridge High School," writes Mr. Rolfe, "devoted all his energies to putting boys into college on strict 'classical' lines. He abolished physics, chemistry, and biology, or (more accurately) reduced them to a lecture a week. The courses for colleges (mostly for neighboring Harvard) were for a five-year period. The whole school was reconstructed for the

purpose of putting boys into college. Scientific apparatus gathered dust!"

There were perhaps a few American high schools that established laboratory courses in chemistry before the one at Cambridge and it is hoped that Mr. Rolfe's letter which I have quoted only in part may bring forth additional information on this subject. The fact that chemical laboratory instruction in high schools was first attempted during the Civil War period and that it afterward underwent a temporary decline seems to have been overlooked by historians of American chemistry.

C. A. BROWNE

U. S. DEPARTMENT OF AGRICULTURE
WASHINGTON, D. C.

In Appreciation of Herbert Freundlich

To the Editor:

Recently, some journals of the learned societies announced the death of Professor Herbert Freundlich, of the University of Minnesota. The announcements were catalogs of his main scientific achievements, records of his academic posts, documentation of his research interests. I wished for an appreciation of his great human qualities, of his honesty and unselfishness, of his genuineness and goodness. An obituary in *Science* did call him one of "science's noblemen." He was that, and more. His character and personality were above the men-of-good-will type, advanced and ideal. All who knew him valued his unassuming attitudes and refreshing friendliness.

Herbert Freundlich's mother was English, and his father German. At times, his use and enunciation of the American language betrayed British residence and ancestry. (He did work and live in England from 1933 to 1938.) When I left his home in Minneapolis one day to take a train back to Chicago, he asked whether I intended to catch a "tram" to the depot. He spoke exceedingly well, with hardly a German accent. His German thoroughness was always evident. I remember his correcting my assertion, minor with respect to the general idea involved, that the German public school system soon after the Franco-Prussian war had three divisions, for children of poor, middle-class, and wealthy parents. He cited his own education in Nassau, and suggested that conditions varied in different parts of Prussia; when he had gone to public school, from 1886 to 1898, conditions had not been so extremely centralized and standardized.

His early scientific work was pioneering, new to the extent of being challenging and audacious. Freundlich's work in colloid chemistry was novel, though in line with the growth of physical chemistry at the beginning of the century, and a logical continuation of his training in Ostwald's laboratory at Leipzig.

He worked with Fritz Haber for seventeen years, from 1916 on. He appreciated this long association, admired the more famous Haber, but he knew his shortcomings and faults. Freundlich's co-workers at the renowned Kaiser Wilhelm Institute for Physical Chemistry in Dahlem, realizing his feeling for Haber and the Institute, gave him as a memento, when he left in 1933, a collection of photographs concerning the laboratories and scientists at the research center. And when he emigrated, he took all his belongings and furniture, astonishing English friends who did not expect him to behave as though the Nazi revolution were irreversible.

I came to know Professor Freundlich when I approached him for material for a book I was writing. "I shall be very pleased to give you any information you need," he replied to my letter of inquiry. The tone was definitely that of an altruist. It was appreciated then; but more now, in retrospect, since others have asked compensation for almost similar material, or have refused to inconvenience themselves.

When I first saw him, I attempted to explain the difficulties in gathering material; the failures, the drawbacks, the expectations, the intentions of my quest heretofore. He dismissed these with a wave of the hand and began to give me the facts I wanted. We talked for several hours that Sunday morning and afternoon, stopping only for dinner and his short after-dinner nap. I had prepared eighty or more questions, and he answered them when he was certain of the information.

Three weeks later I saw him again. Not so much because the matter of acquiring the data was urgent, but more because his magnetic charm was vastly different from anything I had experienced. As a student, I had read about learning at the feet of a great man. That was not understandable as I looked about the university I attended and noticed, for the most part, cold and aloof scholars, disinterested and dispassionate laboratory hermits. Freundlich, on the contrary, was warm and human, wise and personable, knowledgeful and alive. I can understand why the students at the University of Minnesota called him "Uncle Herbert."

Without stopping to rest, he spoke to me for six hours that Saturday evening—remarkable, since he was busy preparing a special set of lectures, expected no compensation, and wanted no publicity. He gave me names and addresses of others who had pertinent data. To one, he later wrote letters on my behalf.

After I had written the first draft of my manuscript, I sent a copy to Professor Freundlich for his evaluation. To my amazement, he went through it with a fine comb, pointing out factual errors, mistaken judgments, and untenable viewpoints. The criticism was not only complete, but it was also unassailable. An expert professional editor could not have done better work.

He was versatile, and broad in his outside interests: literature, music, history, and all phases of science. That was no doubt one reason why he was himself, as he described one woman to me, "exceptionally clever, sensible and, on the whole, free of prejudices." He was himself, as he described Einstein, "a distinctly unselfish man." And he was himself as he described another scientist, "a most decent and courageous man."

M. G.

Out of the Editor's Basket

MOST of us have had at one time or another disastrous experience with rubber stoppers which have become "frozen" to glass tubing, rods, or thermometer stems. R. Szymanowitz, of the Acheson Colloids Corporation, describes a helpful remedy for such trouble. Slip a cork-borer of the right size over the tubing or thermometer stem and force it carefully between the glass and the rubber. This will generally save the stopper as well as the glass.

● We are glad to learn that the Phi Lambda Upsilon chapter at the University of Washington, wishing to promote interest and skill in chemistry in the high schools of the state, presented each of the ten outstanding high schools with a year's subscription to the JOURNAL OF CHEMICAL EDUCATION. We will be glad to give "publicity" in these columns to any other organization which has a similar urge to be professionally philanthropic and would remind them that all such subscriptions are at the special student rate.

● Sigma Epsilon Pi, of the College of William and Mary, awarded subscriptions to the JOURNAL to William R. Hencke and Beverley W. Lewis, cited as "outstanding students in the general chemistry course this year."

● Hugo Zahnd of Brooklyn College sends in the details of a striking experiment performed by Robert Hooke (1633–1703). Five grams of niter, KNO_3, are heated considerably above the melting point in a dry pyrex test tube or porcelain evaporating dish. With great care, the lecturer dusts tiny quantities of powdered charcoal, flowers of sulfur, or powdered red phosphorus into the melt. The non-metals burn brilliantly.

● The creation of a canopy of glaring lights over a city would afford better protection than a blackout in the event of invasion. Such is the opinion of A. F. Dickerson, head of General Electric's illuminating laboratory and an outstanding authority on exterior lighting. It's the same idea as the glaring headlights from automobiles put on a large scale.

This could be done by the installation of numerous small though powerful searchlights, pointed upward and mounted on the roofs of large buildings and atop high elevations. Such a glare of light would tend to blind enemy flyers, prevent them from seeing through the curtain of light and locating vulnerable targets, according to Dickerson.

Such a canopy would also silhouette bombers to the eyes of defending aircraft flying in the sky above the heavy bombers. Such lighting could also be used to decoy enemy pilots by being placed about open fields adjacent to a city, thus extending what would appear the boundaries of the city as viewed from the air.

● We recently received the following statement from the Aluminum Company of America:

"At present, there is no shortage of aluminum for national defense, although civilian applications have been materially curtailed. This reduction in civilian uses has affected many of your readers who have, because of the pressure of defense demands for aluminum, been forced to accept substitutes for the duration of the war. We appreciate the spirit in which they have cooperated with us.

"Just how long we will be able to say 'There is no shortage of aluminum for national defense' is problematical. There may be a shortage next month, in six months, in a year. It depends largely on plane production, not only for the United States but for Britain as well. According to a recent announcement by W. S. Knudsen, the expanded plane program will require an annual production of 1,600,000,000 pounds of aluminum. The best available estimates indicate that this is more aluminum than the whole world produced last year."

● Sulfaguanidine now joins sulfanilamide, sulfapyridine, and sulfathiazole as a powerful aid to the medical profession in combating infectious diseases.

Sulfaguanidine, which is made by the Calco Chemical Division of the American Cyanamid Company, differs from the other sulfa-drugs in its action. Whereas other members of the family are readily absorbed into the blood stream, sulfaguanidine when administered by mouth remains largely in the intestinal tract. Thus the new drug is able to destroy or prevent the growth of certain bacteria in the alimentary canal. This property of sulfaguanidine makes it particularly valuable in combating acute bacillary dysentery, a disease prevalent in armies under unsanitary conditions, and as a prophylactic agent in surgical operations on the colon. By inhibiting the growth of certain bacteria common to the human intestines, sulfaguanidine reduces the possibility of a spread of infection following abdominal operations.

Sulfaguanidine remains in the colon in high concentration and inhibits there the growth of bacteria of the class known as gram-negative bacilli. In this class of bacteria are those which cause acute bacillary dysentery and clinical tests on this disease have shown consistently good, and sometimes sensational, results.

● A paper read recently by Dr. Russell M. Wilder, at the National Nutrition Conference for Defense, contained a number of points worth emphasizing. First of all, there is very convincing evidence that the nation is faced with a serious problem of malnutrition and that despite a so-called surplus of foods a great many of our people are not receiving the fare they need for strength of mind and body. We must mobilize for better nutrition. Government surveys revealed that

in 1936 one-third of all families were buying food which could not provide a diet rating better than "poor" by conservative standards, and that in not more than one family in four would the food provide a diet rated as "good." It seems strange that in spite of our tremendous scientific accomplishments fundamental knowledge in the science of nutrition is of very recent origin. The early findings were vague and negative. It could hardly be expected that men trained in a school of tangible causes and effects would be greatly concerned with the effects of such substances as vitamins when they could only be defined as "things which make you sick when you don't get them." To be sure, doctors have long recognized *severe* deficiency diseases, but when it came to the recognition of the more subtle forms of malnutrition and their relationship with the vitamins they were wary. This is not to be wondered at when one remembers that diet, for centuries, has been a fertile field for quackery. But now we can think and work in terms of *micrograms* of vitamins with chemical names. Henceforth progress should be more rapid.

RECENT BOOKS

ANALYTICAL PROCESSES, A PHYSICO-CHEMICAL INTERPRETATION. *T. B. Smith*, Ph.D., A.R.C.S., B.Sc., Lecturer in Chemistry at the University of Sheffield. Second Edition. Edward Arnold and Co., London, 1940. (Distributed by Longmans, Green and Company, Inc., New York City.) viii + 470 pp. 50 figs. 14 × 22 cm. $5.00.

This book was written as a text for quantitative analysis, stressing the theoretical principles. A few working details are included in this edition but "the chief aim is still to supplement rather than supplant the ordinary type of practical textbook." This in itself makes it unusual.

The author's approach to his subject is unique and has resulted in a most interesting book. Separate chapters are allocated to the precipitations of barium sulfate, lead sulfate, ferric hydroxide, the silver halides, the sulfides, and various separations of mixtures (though mainly calcium and magnesium). Likewise presented are the titrations of the halides with silver nitrate (which includes a discussion of an adsorption indicator), cyanide with silver nitrate, and acid with a base. The theory of oxidation-reduction reactions has been presented in this edition in considerable detail. Less attention has been given to the theory of electroanalysis. Various theories and concepts, such as mass action, ionic product, pH buffers, modern views concerning acids and bases, solubility product, common ion, and colloids are introduced in an early chapter; others are presented in the various discussions as needed.

A rather considerable amount of mathematical physical chemistry is included, which, despite the best efforts, cannot be made really simple, unless the reader is well prepared. For this reason the book is perhaps best adapted to advanced courses or reference use.

REX J. ROBINSON

UNIVERSITY OF WASHINGTON
SEATTLE, WASHINGTON

THE RING INDEX. A LIST OF RING SYSTEMS USED IN ORGANIC CHEMISTRY. *A. M. Patterson*, Professor of Chemistry, Antioch College, and *L. T. Capell*, Associate Editor of *Chemical Abstracts*. American Chemical Society Monograph No. 83. Reinhold Publishing Corp., New York City, 1940. 661 pp. 15 × 23 cm. $8.00.

This work is a compilation of all of the ring systems, about 4000 in number, that have been reported in the literature of organic chemistry through 1939. Only parent ring structures are given, and not those of their derivative compounds. Included are the simple parent rings, parent systems of more than one ring in which the rings have one or more atoms in common, and in deference to current usage, a few ring systems containing ionic valences (*e. g.*, betaine). Cyclic coördination complexes (chelate rings) and systems like biphenyl and diphenylmethane, in which the rings are not joined by actual sharing of atoms, are omitted.

The purpose of the index is to provide a convenient classification of the ring structures used in organic chemistry along with a standard and authoritative system for their numbering and nomenclature. The entries in the index are classified first according to the number of rings present in the structure and then according to the number of atoms in the component rings.

For each entry there are given: (1) a structural formula showing the standard method of numbering the component atoms of the rings; (2) a serial number to serve as a method of identifying the system in references to it; (3) the systematic name of the entry; (4) any generally accepted name which takes precedence over the systematic name (in bold-face type); (5) any other trivial names which are more or less generally accepted (in italics); (6) one or more original literature references, usually including the earliest work giving the structure with certainty; (7) a volume and page reference to the fourth edition of Beilstein's "Handbuch"; (8) other methods of numbering, different from the standard method, but used to an appreciable extent in the literature.

If the name of a given ring system is known, it may easily be located by consulting the alphabetical index at the end of the book, or if the name is unknown, the system may be located by using the general classification which is completely described in the introductory pages.

The authors and their collaborators are to be congratulated on the usable form and the completeness with which THE RING INDEX fulfills its purpose, and it is to be hoped that organic chemists will forego personal preferences and adopt the standard nomenclature and numbering therein set forth.

M. F. ROY

BROWN UNIVERSITY
PROVIDENCE, RHODE ISLAND

HOW TO TEACH A JOB. *R. D. Bundy*, A.M., E.E., Industrial Coördinator, Board of Education, City of Cleveland, National Foremen's Institute, Inc., Deep River, Conn., 1941. 63 pp. 13.3 × 20.2 cm. $1.00.

This manual, intended for foremen, supervisors, job setters, and vocational instructors, has as its purpose the outlining of the steps in the teaching process, and the pointing out of the definite procedures involved. The author looks at the problem in view of the emergency situation brought about through the National Defense program, with the need for training men quickly and efficiently. Considering teaching as "putting across ideas," the author follows the teaching process through the steps of preparation, demonstration and explanation, questions by the teacher, illustration, application, and inspection of the work, ending with a lesson planning analysis. The manual is based on common sense, but it may focus attention upon points often neglected in industrial work.

AN INTRODUCTION TO THE KINETIC THEORY OF GASES. *Sir James Jeans*, O.M., F.R.S. First Edition. The Macmillan Company, New York City; The University Press, Cambridge, 1940. ii + 311 pp. 35 figs. 14 × 22 cm. $3.50.

The purpose of this book, as stated by the author in his preface, is to "provide such knowledge of the Kinetic Theory as is required by the average serious student of physics and chemistry." The book is restricted to that part of kinetic theory which is deducible from the laws of classical mechanics.

A short history of kinetic theory and an interesting discussion of the three states of matter are given in the introductory chapter. This is followed in Chapter Two by an introduction to the main theme of the book. Here the author gives a discussion of the following subjects: the pressure in a gas, equipartition of energy, temperature and thermodynamics, evaluation of Avogadro's and Loschmidts' numbers, specific heats of gases, Maxwell's law, viscosity, conduction of heat, free path, and molecular diameters. Using quite simple mathematics, the author has succeeded very well in giving a physical picture to the phenomena involved. The succeeding portions of the book deal in a more comprehensive fashion with the topics initiated in Chapter Two. The emphasis of subject matter can be seen from the following list of chapter headings: Pressure in a Gas, Collisions and Maxwell's Law, The Free Path in a Gas, Viscosity, Conduction of Heat, Diffusion, General Theory of a Gas Not in a Steady State, General Statistical Mechanics and Thermodynamics, Calorimetry, and Molecular Structure. In the appendix are given Maxwell's proof of the distribution law, a discussion of the H-theorem, normal partition of energy, the law of distribution of coördinates, and also tables for numerical calculation and of integrals involving exponentials.

The book is exceedingly well written and the material presented in a thought-provoking fashion; this, together with the fact that it avoids mathematical proofs of rigorous generality, should make it profitable reading for the average college senior in chemistry and physics.

JOHN R. LACHER

BROWN UNIVERSITY
PROVIDENCE, RHODE ISLAND

SOLUBILITIES OF INORGANIC AND METAL ORGANIC COMPOUNDS. *A. Seidell*, Ph.D., National Institute of Health, Washington, D. C. D. Van Nostrand Co., Inc., New York City, 1940. Third Edition, Volume I. iii + 1698 pp. 15 × 23.5 cm. $12.00.

The final necessity for a complete revision of this compilation has resulted in the present volume, using off-set printing, and the new developments in the microfilm copying of printed pages. It has been decided to include all data upon inorganic and metal organic compounds in this volume, and the results upon the compounds of carbon in a second volume, because of the very large amount of quantitative solubility data available.

As stated in the author's preface, "the old alphabetical plan based on the English names of the compounds has been changed to one having the symbols of the elements as the basis for the alphabetical arrangement. This purely chemical system of presenting the data will be more convenient to chemists having an imperfect knowledge of English, and in addition, makes it possible to use the symbols as guiding marks for locating desired results."

Precision in identifying a compound accurately has been sought by giving both name and chemical formula. Brief remarks in connection with some of the tables are intended "to indicate the general character of the experiments, the methods used, and the probable accuracy of the results. The absence of such remarks may be taken to mean that the determinations presented no exceptional difficulties, that they were made by the usual methods, and with acceptable care."

The variation in source material for the compilation has resulted in the use of different types, and in some defects in the quality of printing, but these deficiencies can be overlooked in view of the fact that this excellent reference book could not otherwise be made available.

PHYSICAL CHEMISTRY FOR STUDENTS OF BIOLOGY AND MEDICINE. *D. I. Hitchcock*, Ph.D., Associate Professor of Physiology in the Yale University School of Medicine. Third Edition. Charles C. Thomas, Springfield, Ill., and Baltimore, Md., 1940. xiii + 264 pp. 22 figs. 14 × 23 cm. $3.50.

The earlier editions were reviewed in J. CHEM. EDUC., **10,** 256 (1933) and **12,** 100 (1935); the earlier criticisms and commendations apply equally to the present volume. It is a most excellent introduction to fundamental physical chemistry concepts and principles which have an important bearing on physiological processes and reactions, but the reviewer feels that the research physiologist and biological chemist should not only have the longer "chemists" course in physical chemistry but in addition, wherever possible, should have an extended following course in thermodynamics with all of the mathematical background which such a course implies.

The present volume represents a rather extensive revision of the second edition. A new chapter has been introduced at the beginning discussing significant figures, logarithms, slide rules, equations and curves, and the nature of thermodynamics.

In Chapter V the Brønsted concept of acids and bases has been considerably expanded. Chapters VII and VIII represent essentially new additions or discussions which have been greatly expanded; we find discussions of ionic mobility and transference numbers including Hittorf's method and the moving boundary method for determining the latter, reversible galvanic cells, reversible and non-polarizable electrodes, concentration cells, liquid junction potentials, and activity coefficients. The chapter dealing with standard electrode potentials and oxidation-reduction potentials has been markedly improved. In the discussion of colloidal systems the new methods of the ultracentrifuge and the Tiselius electrophoretic technic have been introduced. The reviewer feels that the attempt to compress colloid chemistry into slightly over 18 pages of rather large type is not particularly successful, but this criticism is probably equally applicable to the abbreviated treatment characteristic of the other chapters as well.

The volume represents a real revision and not simply a reprinting. Eight new laboratory experiments have been added to the twelve in the second edition. All but one of these deal with electrical technics; the exception shows the dependence of the viscosity of a gelatin sol on pH.

ROSS AIKEN GORTNER

UNIVERSITY OF MINNESOTA
UNIVERSITY FARM
ST. PAUL, MINNESOTA

THE WORLD AND THE ATOM. *C. Møller* and *E. Rasmussen*, University of Copenhagen. Translated by *C. C. Wheeler* and *B. Miall* from the Second Danish Edition. D. Van Nostrand Company, Inc., New York City, 1940. 199 pp. 40 figs. 14 × 22 cm. $2.75.

This small book represents the efforts of a spectroscoper and a theoretical physicist to trace the development of modern atomic physics in non-mathematical language. Such topics as the subatomic particles, radioactivity, the quantum theory and its application by Bohr to the problem of atomic structure, the uncertainty principle, wave mechanics, the atom nucleus, and transmutation are discussed with clarity. It is unfortunate that the translators in many instances have not used the customary English terms instead of translating their Danish equivalents literally. As examples, oscillation number is used in place of frequency and tube meter for Geiger counter. Those who are familiar with the value for Planck's constant in erg-seconds may be confused by its value in electron-volt-seconds especially since the units are not given. Aside from these minor criticisms the book is heartily recommended for collateral reading by college undergraduates in science courses.

JOHN A. TIMM

YALE UNIVERSITY
NEW HAVEN, CONNECTICUT

PHOTOSYNTHESIS. *E. C. C. Baly*, C.B.E., D.Sc., F.R.S., Emeritus Professor of Chemistry in the University of Liverpool. D. Van Nostrand Company, Inc., New York City, 1940. vii + 248 pp. 24 figs. 14 × 22 cm. $4.75.

The contents of this book may be divided into two categories, one a description of research carried out by Baly and his co-workers, the other a discussion of the kinetics and mechanism of photosynthesis.

Baly's experimental investigations related to photosynthesis have been confined to photochemical studies in non-biological systems. These investigations, according to Baly, "led in the end to the photosynthesis in the laboratory of carbohydrates from carbon dioxide and water." The experimental procedures used by the author and his co-workers are described in considerable detail.

Largely because of failure of the various attempts to duplicate his earlier experiments, Baly's results are not at present generally accepted by the majority of investigators in the field of photosynthesis. Nevertheless, the problem of photosynthesis *in vitro* is one which merits much greater attention than it has apparently received in the past.

In the remainder of the book, Baly discusses the process of photosynthesis in plants, and, on the basis of supposed analogies between the natural process and the behavior observed in his *in vitro* studies, proposes a mechanism for the process. In the preface, Baly writes "I am fully conscious that the whole outlook on the problem is a personal one, and for this I beg forgiveness. It is not easy to write impersonally of one's own experiences and at the same time accept full responsibility for their truth."

That his outlook is indeed a personal one is indicated by the fact that in the entire book reference is made to only four publications, excepting those by Baly and co-workers, which have appeared since 1933. The phenomenon of chlorophyll fluorescence receives no mention in the book, despite its apparent importance in the study of photosynthesis.

Individuals doing research in fields related to photosynthesis may find the book interesting because of its rather complete presentation of Baly's experimental procedures and results. However, the book cannot be recommended for students or others wishing to secure a general knowledge of the field, since it fails completely to present a true picture of the present status of the problem of photosynthesis.

WINSTON M. MANNING

UNIVERSITY OF WISCONSIN
MADISON, WISCONSIN

AN ATLAS OF PHYSICAL AND CHEMICAL NOMOGRAMS. *G. V. Vinogradov* and *A. I. Krasilschchikov*. State Publishing House of Technical and Theoretical Literature, Moscow, 1940. 50 roubles.

The alignment charts for physical chemistry contained in this atlas follow the methods of Gibbs, van't Hoff, and other pioneers of physical chemistry who made wide use of geometrical methods for presenting various functional relations as well as for systematic and concise representation of experimental data.

The atlas contains two hundred separate charts covering a great variety of topics in physical chemistry—in particular, the properties of ideal gases and the classical kinetic theory of gases; vapor pressure and boiling points of liquids and solids; adsorption and surface phenomena; dielectric properties and the theory of dipole moments; electrochemistry and the theory of solutions; colloid chemistry and the physics of hydro- and aero-sols. There are fundamental service charts available for conversion of units from one system to another. A detailed text which discusses the general theory of nomograms and gives clear instructions for the use of charts making up the atlas has been appended.

The use of the charts in this atlas, in comparison with the now common arithmetic methods for solving general algebraic equations common in physical chemistry, results in great saving of time, particularly when repeated calculations using the same equation are necessary. Because of this, more extended use of nomographic methods should be encouraged and the appearance of this atlas is very welcome.

Unfortunately both the text and the headings of various charts have been printed only in Russian, which will probably never result in the atlas being of extended practical use in the United States.

From the mathematical point of view, the atlas has been carefully and excellently made up and the reviewer has convinced himself that very accurate results can be obtained from the charts following the methods developed by the authors, which in many cases are original with them.

From the purely technical point of view, however, the makeup of the atlas leaves something to be desired. The quality of the paper, as in most Russian publications, is not very good and it is feared that if the charts are subjected to repeated use their life will be rather brief.

G. B. KISTIAKOWSKY

HARVARD UNIVERSITY
CAMBRIDGE, MASSACHUSETTS

LABORATORY INSTRUCTIONS IN BIOCHEMISTRY. *I. S. Kleiner*, Ph.D., Professor of Biochemistry and Physiology, and *L. B. Dotti*, Ph.D., Assistant Professor of Biochemistry and Physiology, New York Medical College. C. V. Mosby Company, St. Louis, Mo., 1940. 188 pp. 6 figs. 19.5 × 26.5 cm. $1.50.

This manual of biochemical technics substantiates the authors' claim that biochemistry is one of the many sciences needed to provide a better understanding of bacteriology, physiology, pathology, endocrinology, internal medicine—even surgery. The manual's fifteen chapters deal with physical chemistry, carbohydrates, lipids, proteins, tissues, milk, digestion, qualitative examination of blood, qualitative urine analysis, quantitative blood analysis, quantitative urine analysis, metabolism, foods, toxicology, and blood gas analysis. The appendix contains a complete list of the necessary reagents with all the directions for their manufacture, a valuable list of normal blood constituents, as well as the different systems of weights and measures, directions for the use of a centrifuge, a chart of atomic numbers and atomic weights of the elements, and a table of logarithms.

Many of the tests are carried out with a minimum of apparatus and, as the authors state, a good number may be carried out in a doctor's office.

Teachers of biochemistry, doctors, and clinical laboratories will welcome this compact, well-printed, and easily understood manual.

DANIEL RIVKIN

SOUTHWESTERN INSTITUTE OF TECHNOLOGY
WEATHERFORD, OKLAHOMA

A LABORATORY COURSE IN GENERAL CHEMISTRY. *R. H. Crist*, Assistant Professor of Chemistry, Columbia University. First Edition. McGraw-Hill Book Co., Inc., New York City, 1940. xi + 219 pp. 16 figs. 21 × 28 cm. $1.50.

The novel experiments presented constitute a well-organized laboratory course designed specifically for capable students who have presented chemistry for admission to college. "Cookbook" experiments have been avoided and most of the work requires the application of semiquantitative measurements for the interpretation of the observations. An attempt to use this outline without recourse to the recommended weekly conferences would probably not meet with success, as the experiments are by no means simple and even the most able students will require aid in their elucidation. Such students, on the other hand, will feel greater satisfaction in performing laboratory experiments which require their best efforts and in the end will have advanced their knowledge of chemistry to a far greater extent than is possible with the more conventional laboratory outline. For institutions where segregation of students on the basis of aptitude and preparation is practiced, the adoption of this stimulating laboratory manual is recommended for the more advanced group.

LAURENCE S. FOSTER

BROWN UNIVERSITY
PROVIDENCE, RHODE ISLAND

Series of Orthoesters . . .

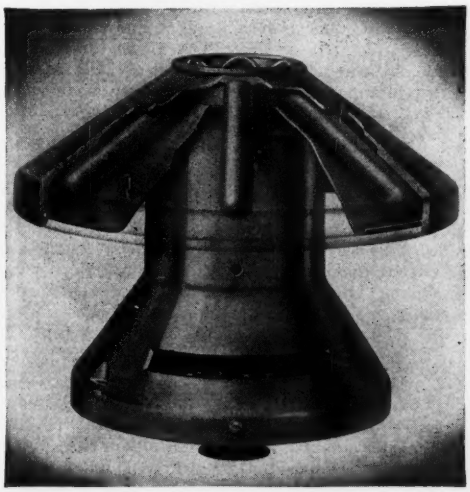
2763	Methyl Orthoacetate	BP 105-108°
2764	Methyl Orthopropionate	BP 123-127°
5094	Methyl Ortho-n-butyrate	BP 144-148°
5101	Methyl Ortho-n-valerate	BP 163-168°
1674	Ethyl Orthoformate	BP 145-147°
P 2851	Ethyl Orthoacetate	Practical
P 2858	Ethyl Orthopropionate	Practical
4257	n-Propyl Orthoformate	BP 92-94°/19 mm.
4244	n-Butyl Orthoformate	BP 122-124°/10 mm.

The ready availability of this group of Eastman Organic Chemicals provides a convenient means of demonstrating in laboratory and lecture work the effect of series of reagents that vary regularly and uniformly. Numerous other similar groups of compounds also are prepared and stocked by Eastman—such as alcohols, amines, mercaptans. Write for a free copy of *Eastman Organic Chemicals, List No. 32,* in which the various series are cataloged, with data. . . . Eastman Kodak Company, *Chemical Sales Division,* Rochester, N. Y.

KODAK

There are more than 3400

EASTMAN ORGANIC CHEMICALS



5000 R.P.M. on D.C.

with underguard No. CT-1055 with six 15 ml. tubes loaded. 4200 R.P.M. on A.C. with same load.

ADAMS ANGLE CENTRIFUGE

These centrifuges offer important advantages over the conventional units. They utilize the new angle principle—the tubes being suspended at a fixed 52° angle—thus, faster sedimentation is achieved by the shorter distance particles are required to travel . . . creating mass, and reaching the bottom more quickly. When at rest, the tubes remain in the angular position and no stirring up of sediment results.

CT-1000 Adams Senior Safety-Head Centrifuge for **Six** 15 ml. **Tubes,** complete with six round bottom brass shields with rubber cushions and three each graduated and ungraduated taper bottom 15 ml. glass tubes. Without Underguard. . . . . **Each $52.50**

CT-1001 Same as the above but without shields or tubes. . . . . **Each $49.50**

CT-1055 Underguard for Safety-Head, made of Aluminum. . . . . **Each $3.50**

Above Centrifuges have universal motors for 110-volt A.C. or D.C. current. They can also be supplied to operate on 220-volt A.C. or D.C. current. Additional charge of $2.00 is made for 220-volt motors.

FOR MICRO and SEMIMICRO WORK—

The above Centrifuge will accommodate six 15 ml. round bottom or taper bottom test tubes. Micro and semimicro tubes (5 ml. to 0.5 ml.) can be accommodated by purchasing extra shields, prices from 50¢ to 75¢ each. No adapters are required.

Other ADAMS CENTRIFUGES and laboratory supplies are described in our new Catalog No. 102CE. If you do not already have a copy write for one on your letterhead, please.

CLAY-ADAMS CO., Inc. 44 East 23rd St. New York, N. Y.

BRYAN VALENCE BLOCKS

FOR TEACHING THE FUNDAMENTAL PRINCIPLES OF THEORETICAL CHEMISTRY

THIS highly instructive set of Valence Blocks, designed by Dr. Arthur Herbert Bryan, Baltimore City College, Baltimore, Maryland, permits the visual presentation of the fundamental principles of theoretical chemistry. The set includes hundreds of differently shaped pieces representing the more common elements and molecules. The pieces are distinctly colored in blue and red, the blue indicating metals and the red indicating non-metals. Valence is indicated by the number of arms extending from each piece. Thus the student is quickly impressed with the harmonious relation of the various atoms and molecules.

These new Valence Blocks, which come packed in an attractive box with complete directions, have 17 specific uses: Demonstrating the construction of numerous chemical formulas and equations; showing the periodic arrangement of the elements; illustrating oxidation and reduction reactions; etc.

Why not add these low-priced, highly educational teaching "tools" to your want list? The cost per set is only..... $1.00

● For down-right economy and all-round satisfaction depend on our catalog No. 45 as your source of chemical laboratory needs. It lists practically everything essential to the proper maintenance of a modern, efficient laboratory.

Chicago Apparatus Company
Scientific Laboratory Equipment
1735 NORTH ASHLAND AVENUE · CHICAGO, U. S. A.

TRADE ANNOUNCEMENTS

Catalog on Thermometers and Hydrometers

A new 28-page catalog on thermometers and hydrometers has just been published by the *George Ruehfel Company, 103 Menahan Street, Brooklyn, N. Y.* This catalog describes the complete line of A.P.I., A.S.T.M., and M.C.A. instruments.

According to the manufacturer, the outstanding feature of this line of hydrometers is the solid alloy ballast that reinforces the tip, thereby greatly reducing the likelihood of breakage.

The catalog gives suggestions for insuring accurate hydrometer readings and has a handy selection chart which lists hydrometers especially calibrated for more than 50 products. Included in the catalog is a double conversion table and a set of tables for converting specific gravity readings into pounds per gallon, °Brix, °Baumé, °Twaddell, °Barkometer or per cent proof spirits, etc.

Catalog is available free if requested on official stationery.

Automatic Check on Distilled Water or Condensate

The checking of distilled water or steam condensate, continuously and automatically, is the function of the DW Checker recently developed by Industrial Instruments. The instrument is based on the electrical conductivity principle. Since the conductivity of distilled water or condensate is a measure of its purity, the readings and control are said to be quite precise. For instance, the presence in distilled water of 1 part per million of chloride ion increases its conductivity about 50%. In the operation of the DW Checker, any increase in conductivity above normal is said to result in the operation of an alarm signal, electric valve or any other desired control equipment. It is further claimed that the instrument panel provides continuous conductivity readings. For use with water stills the meter is calibrated in parts per million of chloride ion in pure distilled water.

The DW Checker is fully AC operated, with a built-in voltage regulator which is said to maintain accurate calibration despite line-voltage variations. The instrument is housed in a steel cabinet measuring $7 \times 12 \times 6\frac{1}{4}$ inches deep, designed for wall mounting. The operating point of the relay, used to control the external alarm or corrective means, may be set for individual requirements by means of an adjustment on the panel. The instrument operates in conjunction with a conductivity cell screwed into a standard connection in the pipe line containing the distilled water, distillate or electrolyte. From that point on the operation of the installation is said to be entirely automatic.

Further information about this instrument may be obtained from *Industrial Instruments, Inc., 156 Culver Avenue, Jersey City, New Jersey.*

Emulsifying Agent

It is claimed that textile oil emulsions which are unaffected by aluminum sulfate, sulfuric acid and other salts and acids can now be made by the use of a new emulsifying agent recently developed by the Glyco Products Co., Inc. This emulsifying agent, known as Emulgor A, is completely described in a bulletin which also gives numerous formulas using Emulgor A.

A copy of the bulletin is available free on request to the *Glyco Products Co., Inc., 230 King Street, Brooklyn, New York.*

Splash Proof Micro Switch

A splash proof, malleable housed, precision switch which can be conveniently mounted from any one of four sides for use as an interlock, limit, or pushbutton switch is announced by the Micro Switch Corporation. It is said to work well where closely held operating points and movement differentials are needed and where any splashing may occur.

The switching element is a Bakelite Micro Switch which provides precision operation, ample overtravel, and long life. The assembly is Underwriters' listed and the switching element alone is also listed by Underwriters' at 1200 watts up to 600 volts A.C.

A removable cover provides easy access to terminals which are supported on the switch element. Conduit measurement is $\frac{1}{2}$". The switch is so mounted inside the housing that it may be easily removed for inspection or change. It is claimed that within the listed rating a life of over a million operations may be expected.

For further information write *Micro Switch Corporation, Freeport, Illinois.*